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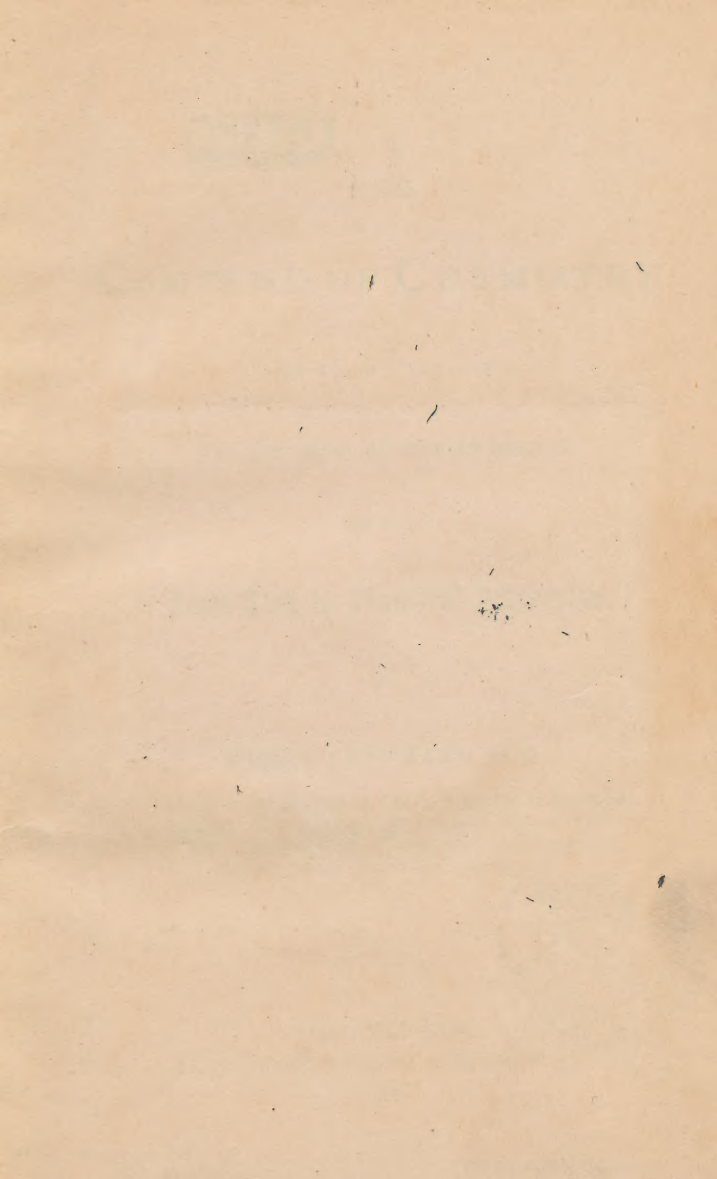
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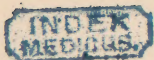
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THE

COMPEND OF CHEMISTRY.

WITH WHICH IS INCLUDED A

SECOND AND REVISED EDITION

—OF THE—

“First Step in Chemical Principles.”

—BY—

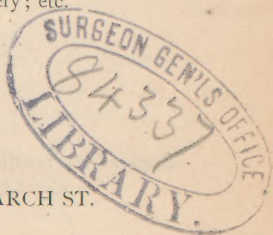
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PREFACE.

THE first twenty-eight of the following pages are a revised reprint of the "FIRST STEP IN CHEMICAL PRINCIPLES," with the success of which book I have every reason to be satisfied, and this fact must constitute the reason for the republication. With regard to the one hundred and twenty-six pages of descriptive chemistry, I have no apology to offer for their publication, nor excuse to make for their character. I am in no manner disposed to follow the custom, now common among writers of chemical manuals, of apologizing for the publication of a book or begging the indulgence of critics. I have put this book out upon the world because it pleased me to do so, and the reviewer who sees in it errors, in fact or method, will but do his duty in pointing them out.

Books like the present are, of necessity, almost wholly compilations, and whatever of merit attaches to the compiler, must be for skill in the selection of matter, perspicuity of expression and symmetry in the treatment of the different topics. The method pursued in this work is the result of experience in the personal instruction of nearly two thousand students.

I am sufficiently conservative to regret the algebraical character which chemistry has assumed of late years, and in the following pages it will be found that this feature has not been brought out to the same extent as in most manuals. I am aware that this may be considered a blemish, but for my part, I think it of far more importance that a student should understand the chemistry of the common proximate principles, than that his mind should be charged with the architectural pictures of molecules, now so popular. I have ventured in many instances to write reactions in which atoms are set free in uneven numbers, and I consider it sufficient defence to say that Bloxam does the same.

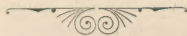
The index will, I believe, make up for disconnections in the description, etc.

H. L.

715 Walnut Street, Philadelphia,
September, 1881.



“I HAVE gathered a posie of other men’s flowers, and nothing
but the thread that binds them is mine own.”



COMPEND OF CHEMISTRY.

I. ATOMS and MOLECULES.

THE material objects around us present an infinite variety of color, form and general qualities. We can, by mechanical action, reduce most of them to fragments, or even to fine powder. The Science of Chemistry has taught us that most bodies are capable of another kind of division, which consists in separating the body into two or more substances unlike each other, and unlike the original substance. This change is called DECOMPOSITION. A fragment of common salt, for instance, may, by mechanical action, be converted into a fine powder, or it may be decomposed into two substances; one, a brilliant metal, and the other a greenish colored gas. These two substances are incapable of further decomposition, and are called ELEMENTS. Chemists, by experimenting upon a great variety of bodies, have established the existence of sixty-five elements, to each one of which a name has been given. A certain force, or form of energy, known as CHEMICAL AFFINITY, causes these elements to unite with one another and form compounds. Some of the elements occur in nature in the free or uncombined condition, but most of them are in combination, hence the various objects of chemical study are either *elementary* or *compound*. The elements are usually arranged in two groups or classes, METALS and NON-METALS; the METALS are generally distinguished by a high lustre when polished, by insolubility in water, and by special powers of combining with acids. The distinction, however, between the two groups, is very vague, and the arrangement will doubtless soon be abandoned. Many of the elements are quite rare, and need not be described in an elementary treatise. The following table contains the more important ones—a complete table is given in the appendix.

1st Group, so-called non-metals.

Oxygen,	Chlorine,	Nitrogen,	Carbon,
Sulphur,	Bromine,	Phosphorus,	Boron,
Selenium,	Iodine,	Arsenic,	Silicon,
Tellurium,	Fluorine,	Antimony.	

2nd Group, so-called metals.

Potassium,	Barium,	Aluminium,	Tin,
Sodium,	Strontium,	Iron,	Copper;
Silver,	Calcium.	Manganese,	Mercury,
Lithium,	Zinc,	Chromium,	
Hydrogen,	Magnesium,	Nickel,	Bismuth,
Lead,	Cadmium,	Cobalt,	Gold.

The arrangement in sub-groups is a rude classification of the elements according to similarity of properties, but it is by no means perfect.

Atomic Theory. If we reduce a piece of sulphur, or any other element to powder, this under the microscope will be seen to be capable of being furthur divided, and it would seem as if no limit existed to this division. Chemists, however, are now generally of the opinion that a limit does exist, and that by continued division we would reach particles of infinite hardness, incapable of diminution or destruction, and of being scratched or broken. These have never been obtained. The belief of their existence is dependent upon some very difficult mathematical and physical considerations; they are infinitely small, and equally hard no matter what the nature of the mass which they make up. These particles are called **ATOMS** (a word derived from the Greek, and signifying indivisible), and any mass of elementary matter consists of a collection of a greater or less number of these atoms. When, therefore, we divide the piece of sulphur, we merely separate the different atoms from one another. Sulphur is soft; carbon in the form of the diamond is very hard; but this difference in hardness between the two is due simply to the firmness with which the atoms of the two bodies hold together. The atoms of the diamond have a strong affinity, those of sulphur a weak affinity; but, in each body, the atoms themselves are of infinite hardness, resisting all wear and tear.

It is now a generally accepted theory that even in elementary bodies the atoms are not perfectly free but associated in pairs. A combination of atoms is called a **MOLECULE**.

Atoms of the same kind are held together by **COHESION**; atoms of unlike kinds are held together by **CHEMICAL AFFINITY**.

When a solid becomes a liquid, or a liquid becomes a gas, or the reverse of these changes occurs, the molecules are not changed, but merely separated from one another. Hence the atoms in sulphur vapor are as hard and solid as those of solid sulphur, only in the vapor the pairs or molecules which they form are separated by greater distances than in the case of the solid. The following will render this point clear :

aa aa aa aa	The element in the solid state.
aa aa aa aa	“ “ liquid “
aa aa aa aa	“ “ gaseous “

This does not represent the proportionate separation, but only the general idea that this change of state is a simple separation of molecules. According to this view, bodies in passing from solid to liquid, and from liquid to gas, should increase in volume, and in the reverse processes should decrease; and this we find to be the case with the vast majority of substances.

Atomic Weights. Chemists have never been able to isolate or render visible atoms or molecules. Their size and weight remain entirely a matter of speculation and theory. Nevertheless, the progress of chemical research, and the application of mathematics, have rendered probable some general principles. These are:—

1st. That the atoms of each element have a constant and definite weight.

2nd. That the atom of hydrogen is the lightest of all.

3rd. That combination takes place among atoms under the influence of cohesion or chemical affinity.

Starting with the first two principles, a series of numbers has been obtained which represent the weight of each atom compared to the atom of hydrogen. These numbers are called **ATOMIC WEIGHTS**. A complete table of them is given in the appendix. It is not necessary for the student to commit them to memory.

2. NOTATION.

A **CHEMICAL** symbol is an abbreviation of the name of an element; in most cases an initial letter is used, as C for carbon; P for phosphorus. As a number of the elements have names beginning with the same letter, proper distinction is obtained by assigning the single letter to the most common, and attaching small letters to the other initials. Thus, C stands for carbon, Ca for calcium, Cl for chlorine, Cd for cadmium. Certain elements have different names in different languages, and for these the symbol is formed from the Latin name. Iron, for instance, is represented by Fe, (*ferrum*); lead by Pb, (*plumbum*); silver by Ag (*argentum*); potassium by K (*kalium*).

Symbols are absolutely invariable. No symbol represents two elements, no element has more than one symbol. The student should commit to memory thoroughly and accurately the symbols

of all the important elements. A complete table of symbols will be found with the table of atomic weights.

To express combination between elements, in other words to express the composition of a compound body or the molecules of an elementary body, the symbols are to be written together like the letters of a word. Such a collection of symbols is called a FORMULA.

The symbol, however, not only represents the element, but also one indivisible particle of it, that is, one atom. Hence, the expression CaO not only shows a compound consisting of calcium and oxygen, but also indicates that it contains a single atom of each element. CaO_2 shows that two atoms of oxygen are present, and one of calcium. In writing these expressions certain rules are followed :

1st. To multiply any single atom, a small number is attached to the lower right hand, as seen above where O_2 indicates two of oxygen. The formula $\text{C}_2\text{H}_4\text{O}_2$ shows a combination consisting of two atoms of carbon, four of hydrogen, and two of oxygen.

2nd. To multiply several atoms by the same number, we put a large figure in front. Thus 2HClO is equal to $\text{H}_2\text{Cl}_2\text{O}_2$; that is, the large figure multiplies the whole expression. This rule gives much trouble to beginners.

3rd. To multiply a portion of an expression, several methods are in use. We may enclose the part to be multiplied in a parenthesis and attach the proper number to the lower right-hand corner. $\text{Ba}(\text{NO}_3)_2$, for instance, equals BaN_2O_6 ; $\text{C}_6\text{H}_8(\text{NO}_2)_2\text{O}_5$ equals $\text{C}_6\text{H}_8\text{N}_4\text{O}_9$. The effect of the small figure is limited to the part within the parenthesis. This method is especially adapted to multiplying symbols in the middle or at the end of a formula. To multiply the symbols at the beginning of a formula, we usually point off or punctuate the part to be affected, and place a large figure in front. Some irregularity prevails as to the particular sign used, the comma and semicolon both being employed. It is sufficient for the student to bear in mind that a punctuation mark or plus sign occurring in a formula will stop the multiplying effect of the large figure at the beginning of the expression. For instance, $2\text{C}_2\text{H}_5\text{H}_2\text{N}$ is equal to $\text{C}_4\text{H}_{10}\text{H}_2\text{N}$; similarly, in $2\text{FeSO}_4 + \text{HCl}$, the letters following the plus sign are not affected by the figure 2. If we wish to carry the multiplying effect to the end of the expression, we enclose it in parenthesis, thus, $2(\text{FeSO}_4 + \text{HCl})$. Here all the letters are equally influenced.

Since the symbol of each element represents one atom, it follows that every symbol carries with it an idea of quantity. If we write HCl , the meaning is not merely that hydrogen and chlo-

rine are in combination, but that the amounts by weight are in the proportion of the atomic weights; *i. e.*, 1 (atomic weight H) to 35.4 (atomic weight Cl). When the symbol is multiplied, the weight is also multiplied. For instance, H_2O represents 2 parts by weight of H to 16 of O; $HgCl_2$ represents 200 parts of mercury and 70.8 (35.4×2) parts of chlorine. From these examples it appears that formulæ give no exact idea of the percentage of the different elements, unless we multiply each symbol by its atomic weight. The following illustration will perhaps make this point clearer. By burning sulphur in air, or in oxygen, we obtain a gas which contains equal parts by weight of S and O. To use the formula SO would be incorrect, for the table of atomic weight shows that the atom of S equals 32, while that of O equals 16; SO, therefore, would show a relation of 32 to 16. The relation is 32 to 32, that is, equal parts; hence, we must use the formula SO_2 , which gives us $S = 32$; $O_2 = 32$ (16×2), the proper proportion. In the same way, we may reduce a more complicated formula. Potassium carbonate is written K_2CO_3 . Referring to the table of atomic weights, we find the following numbers:

$$\begin{array}{lll} K = 39 & \text{consequently} & K_2 = 78. \\ C = 12 & & " & C = 12. \\ O = 16 & & " & O_3 = 48. \end{array}$$

The sum of the atomic weights is called the MOLECULAR WEIGHT. In the example above given, the sum is 138; we cannot reduce this sum in any way, except by taking away one of the atoms; for, by the atomic theory, we cannot remove a portion of an atom. To take away any atom is to change the composition of the body; it is no longer potassium carbonate. Therefore, we say that the smallest portion of potassium carbonate that can exist will be 138 times as heavy as one atom of hydrogen. The same is true of the molecular weight of any body. It will represent the relation between the weight of the smallest possible quantity of the body, its molecule, and one atom of hydrogen. With a number of compound gases, it has been found that the molecular weight is equal to twice the specific gravity compared to hydrogen; and the inferences from this fact will be discussed later.

3. NOMENCLATURE.

THE names of chemical compounds are regulated by a system which depends essentially upon the employment of certain terminations.

The names of most of the metals end in "*um*." When the

metal combines with oxygen the name of the compound is formed in some cases by changing "um" into "a." Potassium (K), for instance, becomes, by taking oxygen, potassa (K_2O); strontium (Sr) becomes strontia (SrO); magnesium (Mg) becomes magnesia (MgO). This rule is of limited and uncertain application.

Chemical compounds which contain only two elements are called BINARY COMPOUNDS. They are usually named by joining the names of the elements present, and attaching to the non-metallic one, when such is present, the termination of "IDE." This termination may be conveniently regarded as an equivalent of the phrase "nothing else"; that is, wherever it occurs it indicates that nothing else is present except what is expressly mentioned. Potassium *iodide*, for instance, can contain nothing else but potassium and iodine; copper *sulphide*, can contain nothing but copper and sulphur.

PbO ,	Lead Oxide.
$NaCl$,	Sodium Chloride.
$AgBr$,	Silver Bromide.

When both the elements are metals, or both non-metals, the system is somewhat uncertain, and in many cases the name is left to individual fancy. Thus the compound of sulphur and chlorine is called by some chlorine sulphide, by others sulphur chloride.

In many books, especially in older works, the word "of" will be found frequently used in the names of compounds. Instead of copper sulphide, we may see sulphide of copper; iodide of potassium for potassium iodide. This system was introduced into chemistry by an original mistranslation of French phrases in which the word "*de*" occurred. Good usage has fortunately justified the entire omission of the word in English chemical nomenclature.

As elements may combine in several proportions, forming several different compounds, this termination *ide* does not suffice for proper distinction, and chemists employ a system of prefixes attached to the name of the non-metallic element. These prefixes are formed mostly from the Greek or Latin numerals. The bodies Cu_2O and CuO are both properly called copper oxide, because they contain only copper and oxygen; but they are quite different substances. In the same way, SO_2 and SO_3 are both sulphur oxides, but must be distinguished from each other. The distinction between the bodies just mentioned is in this way:

Cu_2O ,	Copper <i>suboxide</i> .
CuO ,	" <i>monoxide</i> , (also <i>protoxide</i> .)
SO_2 ,	Sulphur <i>dioxide</i> , (also <i>deut</i> or <i>binoxide</i> .)
SO_3 ,	" <i>trioxide</i> , (also <i>teroxide</i> .)

The terms in parenthesis are now rarely used. In regard to the first prefix *sub*, it is of rather uncertain meaning. It generally indicates deficiency; that is, that the quantity of the element to which it is attached is less than it should be. We can apply the term *sub* to most compounds in which the number of atoms of the metal exceeds that of the other body. Hence, Pb_2Cl , Zn_3I_2 , Cu_4O , would be sub-compounds.

With some of the elements, however, the *normal* proportion of combination is two of the first to one of the second, and here it has become universal custom to use the prefix *mon*. We have, therefore,

K_2O ,	Potassium <i>monoxide</i> ,
Na_2O ,	Sodium <i>monoxide</i> .
Ag_2O ,	Silver <i>monoxide</i> .
K_2S ,	Potassium <i>monosulphide</i> ,

and so on with all the series. The following includes all the elements which show this exception :

NON-METALS.

Chlorine,
Bromine,
Iodine,

METALS.

Hydrogen,
Potassium,
Sodium,
Lithium,
Silver.

It is especially in the combinations of these bodies with oxygen and sulphur, that the irregularity is shown.

Some elements form compounds in which the proportion is as 1 to $1\frac{1}{2}$, but as fractions are not allowed in formulae, the whole expression is multiplied by 2, which gives the proportion 2 to 3. $\text{FeO}_{1\frac{1}{2}}$ becomes, therefore, Fe_2O_3 . These are called *sesqui*-compounds, and the above expression is iron *sesqui*-oxide. The word *sesqui* means one and a half and conveys the idea that the relation between the metal and the other element is as 1 to $1\frac{1}{2}$ (2 to 3). Higher proportions also occur. We have, for instance,

CCl_4 ,	Carbon tetrachloride or quadrichloride.
PCl_5 ,	Phosphorus pentachloride.

In assigning names to compounds containing more than two elements, a great difficulty occurs from the very large number of such compounds which may be formed. To express the names of all the elements, would in most cases make very long words, so the general custom is, either to leave some elements unexpressed, or to give a name to a group of two or more of the elements. For instance, the substance KHO is called potassium hydrate. In this name the K and H only are indicated; the O is understood to be present. The compound KCN is called potassium cyanide, the combination CN being called cyanogen.

This latter method of giving special names to groups is very common in organic chemistry.

Among the most important of the compounds containing three elements are those which we call SALTS. A salt may be defined as the result of the action of metal or oxide of metal upon an acid. If we put zinc or zinc oxide into sulphuric acid, we get a zinc salt; in this case a zinc sulphate. We can get salts also by direct union of many oxides; for instance, when calcium oxide, CaO , acts upon carbon dioxide CO_2 , we get calcium carbonate, CaCO_3 , which is a salt.

Most salts contain *three* elements of which, *oxygen* is one, and the names are made by joining the names of the other two elements and adding to the non-metallic one certain syllables which not only indicate the presence of oxygen, but also partly the amount. These syllables are "ATE" and "ITE." The former indicates the greater quantity oxygen. Thus potassium sulphate and potassium sulphite both contain oxygen, but the former, (sulphate) contains the more oxygen. Sodium nitrate and sodium nitrite contain the same elements, but their composition is KNO_3 and KNO_2 , respectively.

It will aid in the comprehension of this subject if we extend the principle which has already been mentioned when speaking of the termination "IDE." It was pointed out, that this syllable could be regarded as equivalent to the phrase "nothing else." In the same manner, the syllables "ATE" and "ITE" are to be regarded as meaning "something else," and that something else is generally oxygen. With these points in mind, the student will recognize, at a glance, that while in sodium sulphide but two elements are present, sodium sulphate and sulphite will contain three.

These two terminations are not sufficient to distinguish all the salts that may be formed from certain elements. For instance, potassium, chlorine and oxygen will unite in four different proportions, forming KClO_4 , KClO_3 , KClO_2 , KClO . In such cases the important or most common compound is distinguished by the termination "ATE," and the one containing the next lower amount of oxygen by the termination "ITE."

The other compounds are indicated by the use of certain extra syllables, "HYPO" and "HYPER," the latter now generally abbreviated to "PER." The significance and use of these syllables are shown below :

KClO_4	Potassium perchlorate.
KClO_3	" chlorate.
KClO_2	* chlorite.
KClO	" hypochlorite.

From this table it is seen that "PER" intensifies the meaning of any termination, that is, indicates more oxygen than if the termination were used alone; while "HYPO" diminishes the power of a termination, that is, indicates a smaller amount of oxygen than would be present if "HYPO" were not used. Several other series of salts show the same principle, although not so perfectly as that above given.

Na_2SO_4	Sodium sulphate.
Na_2SO_3	" sulphite.
Na_2SO_2	" hyposulphite.
Ag_3PO_4	Silver phosphate.
Ag_2PHO_3	" phosphite.
AgPH_2O_2	" hypophosphite.

When, in such compounds, hydrogen is present, we might use a similar system; but the custom of chemists has decreed that a different method shall be adopted.

Taking the series of chlorine compounds given above, in place of the potassium salts we might have HClO_4 , HClO_3 , HClO_2 , HClO , and these might be called hydrogen perchlorate, hydrogen chlorate, etc.; such names are used by a few persons, but have not become current. The more frequent method is to drop the word hydrogen, change the termination ATE into IC, the termination ITE into OUS and add the word ACID.

The series would therefore be :

HClO_4	Perchloric acid.
HClO_3	Chloric acid.
HClO_2	Chlorous acid.
HClO	Hypochlorous acid.

Note particularly that the *prefixes* are retained without change and that the syllable IC is found whenever, in the metallic salt, the termination was ATE, and the syllable OUS is found when the name has come from a compound ending in ITE.

In the same way we have the following transformations :

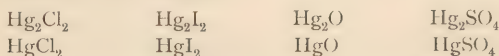
Potassium sulphate		Sulphuric acid.
K_2SO_4	corresponds to	H_2SO_4
Potassium sulphite		Sulphurous acid.
K_2SO_3	"	H_2SO_3
Potassium hyposulphite		Hyposulphurous acid.
K_2SO_2	"	H_2SO_2

Strictly speaking, no necessity exists for this variation. The

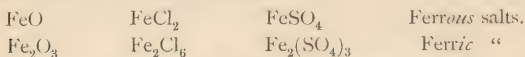
compounds containing hydrogen ought to be regarded as salts. They exhibit, however, some incidental properties which distinguish them from the rest of the salts. They have a sour taste, redden vegetable blues, and have, as a rule, a wider range of chemical action. They stand out as a group; and from a very remote period have been called acids. The term is too well established in chemistry to be set aside. Sometimes we have bodies in which the hydrogen is only partly replaced by another metal, and such bodies are intermediate between the acids and the salts. Thus, KHSO_4 is at once a potassium and a hydrogen compound. In such cases the name is a combination of both systems. The above compound, KHSO_4 , is called acid potassium sulphate. Here the word *acid* calls attention to the existence of hydrogen, while the rest of the elements are indicated by the latter part of the expression.

These acid salts are not unfrequently called *bi*-salts. Acid potassium sulphate, for instance, is generally known in commerce as potassium bisulphate; the corresponding acid carbonate, KHCO_3 , as bicarbonate. This use of the syllable *bi* is improper. If it means anything in this connection, it is that two molecules of acid are present, which is not the case. In a few compounds of exceptional composition the title is used for want of a better one. $\text{K}_2\text{CrO}_4\text{CrO}_3$, for instance, is called potassium bichromate. It is not properly so-called; it does not contain two molecules of chromic acid, as CrO_3 is chromium teroxide, but the more scientific title, anhydrochromate, will not be likely to find favor, and the incorrect name will long be used.

Of late years, the termination *ous* and *ic* have been applied to metallic compounds in which the metal forms two sets of compounds containing the same element, but in different proportions. Mercury forms two chlorides, two iodides, two sulphides, etc. The two series are as follows:



The bodies in the upper row are called mercurous salts, those in the lower row mercuric salts.



Note particularly how these terminations are applied. They indicate not the amount of the metal, but of the other substance; *ous*, as usual, means less than *ic*.

4. LAWS OF CHEMICAL COMBINATION.

CHEMISTRY is an exact science, that is, its laws and principles are mostly known and established. The great law of Chemistry is the LAW OF CONSTANT PROPORTION and is generally expressed thus :

All chemical compounds are definite in their nature, the ratio of their constituents being constant.

In accordance with this we find that any substance in the pure state always contains the same elements and in the same proportions. It must not be supposed that the above law indicates that the elements can only combine in one proportion; on the contrary, some of them, carbon and hydrogen, for instance, combine in very many proportions; but the important point is that each of these compounds is a different substance. As an excellent illustration of these principles, the compounds of mercury and chlorine may be mentioned. If we unite 200 grains of mercury with 35.4 grains of chlorine, we get calomel; if we unite 200 grains of mercury with 70.8 grains of chlorine, we get corrosive sublimate. The two bodies are so different that if we did not know their composition, we would suppose them to contain different elements. Calomel is insoluble in water, and non-poisonous. Corrosive sublimate is soluble in water, and one of the most violent of poisons. When we examine a great variety of chemical compounds, we notice that even when two elements combine in different proportions, a simple arithmetical relation exists. A few cases will make this clear :

Two compounds of oxygen and carbon are known, CO and CO_2 . Two compounds of sulphur and oxygen are known, SO_2 and SO_3 . In the series of compounds of potassium, chlorine and oxygen, given a page or two back, we have



in which the proportion of oxygen is seen to increase regularly. The observation of many facts of this kind has given rise to a sort of principle in chemistry, which is called the LAW OF MULTIPLE PROPORTION; expressed thus : *When two bodies combine in more than one proportion, the higher proportions are multiples of the lower by a whole number.*

This law amounts to saying that if two elements are in combination in a low proportion, and we wish to add more of one of them, we must add either one, two, three etc. times as much. This principle, however is not a law; it is true of the simpler

bodies only ; quite a number of elements combine in almost every proportion and in many cases the multiple relation is only apparent, having been obtained by a sort of arithmetical trick. The great and important law of chemistry is that of constant proportion.

The atomic weights express the proportion in which the elements combine. The atomic weight of hydrogen, for instance, is 1 ; that of bromine 80, and the only known compound of these elements contains 80 parts of bromine, to one part of hydrogen. In some elements, the power of combination is such that two atomic weights of the one are required to one atomic weight of the other ; and with other elements, the proportion is three to one, and even higher. Thus, the atomic weight of oxygen is 16, but the compound of H and O is 2 parts of H to 16 parts of O, and its formula must therefore be H_2O . The atomic weight of nitrogen is 14, but the compound of N and H is 14 parts of N and 3 parts of H, making NH_3 . Compound bodies combine in proportions which are either the sum of the atomic weights of their elements, or are in some simple relation thereto. Lime and carbon dioxide combine and form chalk, calcium carbonate. If we wish to find the proportion by weight, we proceed thus : lime is CaO , carbon dioxide is CO_2 :

$$\begin{array}{rcl} \text{atomic weight Ca} & = & 40 \\ \text{“ “ O} & = & 16 \end{array}$$

Molecular weight, 56

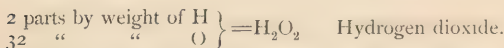
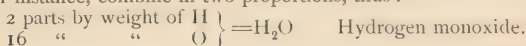
$$\begin{array}{rcl} \text{atomic weight C} & = & 12 \\ \text{“ “ O}_2 & = & 32 \end{array}$$

Molecular weight, 44

The proportion is 56 grains of lime to 44 grains of carbon dioxide.

5. ATOMICITY.

IT HAS already been mentioned that elements may combine in several proportions, producing in each case distinct bodies. When compounds containing the same elements are compared, we generally find one proportion which seems to be the most natural one ; it is either most frequently or easily produced, or it is the one least liable to change. Hydrogen and oxygen, for instance, combine in two proportions, thus :



These two bodies are very different. The first is *water*. It is

formed whenever hydrogen is allowed to burn in air; and it is well known as a compound not liable to decompose. The second substance is very difficult to prepare and to preserve; it is liable to explode. We can have no doubt, therefore, that the natural proportion of combination between H and O is H_2O . Carbon forms with oxygen two well marked compounds, CO and CO_2 . CO is formed when carbon is burned in a *deficient* supply of air, but CO_2 is formed when the carbon burns under natural conditions, in a free draft of air or oxygen. CO, besides, shows a tendency to take upon more oxygen, especially when heated, and it will combine with chlorine even at ordinary temperatures. CO_2 , on the other hand, shows no tendency whatever to combine with oxygen or chlorine. This list might be continued at great length, for all the elements have been more or less completely examined with reference to this point.

In developing this principle it has been found convenient to take the atom of hydrogen as a point of comparison, and to arrange each element according to the number of hydrogen atoms with which it forms the most permanent combination. Taking some important elements, for instance, we find their compounds with hydrogen as follows:

Cl	combines with one H	forming HCl.
Br	“ “ “ H	“ HBr.
O	“ “ two H	“ H_2O .
S	“ “ “ H	“ H_2S .
N	“ “ three H	“ H_3N .
As	“ “ “ H	“ H_3As .
C	“ “ four H	“ H_4C .

These are not the only compounds that can be formed from these elements, but they are those which do not show a tendency either to take new atoms, or give up what they already possess.

If an element does not form a compound with hydrogen, we may either combine it with some other body, and then compare that to hydrogen, or we may displace hydrogen from combination and thus get an idea of the number of hydrogen atoms to which the element is equal.

The greatest number of hydrogen atoms with which any element combines is called its **ATOMICITY** or **QUANTIVALENCE**. To each degree of atomicity a name has been given; thus,

If one atom of the body combines with one atom of H it is called a *monad*; if one atom of the element combines with two atoms of H, it is called a *dyad*, and so on. We have, thus, triads, tetrads, pentads, hexads, and even higher terms. In order

to indicate atomicity, small marks are attached to the *upper* right hand corner of the symbols. The marks are the Roman numerals.

The marks are not much used in formulæ as they produce confusion.

In place of the terms just given, some writers use monatomic or univalent, for monad ; diatomic or bivalent, for dyad ; triatomic or trivalent, for triad ; etc. These have no advantage over the terms used in the text.

Atomicity has nothing to do with the energy or activity of the element. It is a measure of capacity only. Bodies of high atomicity are often of weak affinity, while some of the strongest chemical agents are of low atomicity. Chlorine has only one-third the atomicity of nitrogen, but it is many times more energetic as a chemical substance.

A knowledge of the atomicity of the elements is a key to the composition of all their important and more permanent compounds. Accordingly, the following table should be thoroughly committed to memory, since it will enable the student to write correctly the formulæ of many common chemical substances.

Atomicities of the more important elements :

Monad.	Dyad.	Triad.	Tetrad.
Cl	O	N	C
Br	S	P	Si
I	Ba	As	
F	Ca	Sb	
H	Sr	Bi	
K	All the common metals except silver and hydrogen.		
Na			
Ag			

When elements are combined in such proportion that their atomicities are equal, the compounds are said to be SATURATED. This use of the word must be carefully distinguished from its older and more common use, meaning that a body has dissolved or absorbed as much of any substance as it can take up. In this sense we speak of *saturated solutions*, meaning solutions which contain as much of any substance as can be dissolved ; we speak also of gases being saturated with moisture, meaning that they contain as much moisture as can be held by them under the conditions.

The above table can be used not only as a guide to the compounds which each element forms with hydrogen, but also as a guide to the compounds which the elements form with each other. The system is quite simple. Taking the monad group, for instance, it will be at once understood that, as they are all equal to

one atom of H, they are equal to each other. Hence, K and Cl will combine in equal atoms forming KCl, potassium chloride. Similarly we will have NaBr, AgI, etc. The elements in the second column have twice the combining capacity of those in the first column ; we will find, therefore, that the compound of sodium and oxygen will be Na_2O .

The general rule is that the elements unite in such proportions that the degrees of atomicities are equalized. Suppose we have a compound of Cl and Sb ; Sb is a triad, that is equal to three hydrogen atoms, while Cl is a monad and equals only one hydrogen atom. It will therefore take 3Cl to have the same capacity as one Sb, and the proper formula will be SbCl_3 . If a compound of C and Cl be desired, we find that as one is a tetrad and the other a monad, the resulting compound will be CCl_4 .

The most difficult application of the rule is where triads and dyads are united, as in the compounds of oxygen and sulphur with nitrogen, phosphorus, arsenic and antimony. A few trials will show that to make the two atomicities equal we must multiply the dyad by 3 and the triad atom by 2. For example, Sb_2S_3 must be the formula of antimony sulphide, for :

$$\text{Sb}_2 = 3 \times 2 = 6.$$

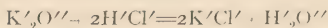
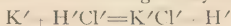
$$\text{S}_3 = 2 \times 3 = 6.$$

A few formulæ are here appended as additional illustrations :

Monad with monad	$\text{H}'\text{Cl}'$	$\text{K}'\text{Cl}'$	$\text{Na}'\text{Br}'$
" " dyad	$\text{H}''_2\text{O}''$	$\text{K}''_2\text{O}''$	$\text{Cu}'\text{Cl}'_2$
" " triad	$\text{H}'''_3\text{P}'''$	$\text{Ag}'''_3\text{Sb}'''$	
" " tetrad	$\text{H}^{iv}_4\text{C}^{iv}$		
Dyad " dyad	$\text{Cu}''_2\text{O}''$	$\text{Zn}''\text{S}''$	$\text{Fe}''\text{O}''$
" " triad	$\text{P}'''_2\text{O}'''_3$	$\text{N}'''_2\text{O}'''_3$	$\text{As}'''_2\text{S}'''_3$
" " tetrad	$\text{C}^{iv}\text{O}^{iv}_2$	$\text{C}^{iv}\text{S}^{iv}_2$	

It must, however, be borne in mind that compounds exist in which the proportion of the atoms differs from what the rule requires ; but the essential character of such compounds is, a tendency to change, either by taking new atoms or giving up some that they already possess. The normal compound of oxygen and carbon is, of course, CO_2 , in which the one tetrad C atom is exactly saturated by the two atoms of dyad O. We are, however, acquainted with a well defined body having the formula CO. This substance is rather a proof of the rule than an exception to it, for it shows a strong tendency to take up other atoms in order to complete its structure. Heated in the air, it combines with oxygen and forms CO_2 , and it unites with chlorine, forming a definite compound which will, of course, have the formula COCl_2 since it requires two chlorine atoms to perform the function of one oxygen atom.

So, also, the student must not fall into the error of supposing that bodies in which the degrees of atomicity are equalized, are necessarily without chemical activity. On the contrary, some of our most active chemical agents are saturated compounds. Hydrochloric acid, HCl , is an example of such a substance. It shows no tendency to take new atoms except under the condition that it at first give up a portion of its structure. It will dissolve potassium, for instance, but only by first losing its hydrogen, into the place of which the potassium enters. It will dissolve potassa (potassium oxide K_2O), but only by exchanging its hydrogen for the potassium. The changes may be represented as follows :



This is, in fact, the nature of common chemical changes : they are substitutions of one element for another, the change always taking place in such a way that the element driven out is exactly equal in atomicity to the one that enters the combination. If, instead of acting on hydrochloric acid with potassium, we use zinc, the quantity of HCl will have to be increased; the reaction $\text{Zn} + \text{HCl}$ cannot take place, since one atom of zinc must drive out two atoms of hydrogen, zinc being a dyad. Therefore, we say $\text{Zn} + \text{H}_2\text{Cl}_2 = \text{ZnCl}_2 + \text{H}_2$, which is in strict accordance with fact, as showing that one atom of zinc will set free twice as much hydrogen as one atom of potassium will.

The degrees of atomicity given above are not invariable. The circumstances under which the variation takes place cannot be very well defined, but the extent or rate of variation is governed by a very simple law, to which only a very few exceptions need be made. *When an element changes its atomicity, either increasing or diminishing, the change is by two degrees at a time.* Elements of even atomicity remain even, passing, for instance, from hexads to tetrads and finally to dyads, or *vice versa* : elements of uneven atomicity remain uneven, passing from pentads to triads and monads.

Certain elements vary in atomicity in a way that appears to be exceptional, but in which we can, by a simple and not improbable supposition, account for the change and yet preserve the application of the law. The elements referred to are iron, manganese, chromium, aluminium, and a few other less important metals, and, among the non-metallic bodies, carbon. These have the property of combining with themselves in such a manner as to form double atoms, possessing an atomicity greater than either atom singly, but less than the sum of the atomicities of the two atoms. For instance, iron, which is generally a dyad, becomes in certain compounds a tetrad, but instead of forming compounds

upon this basis, two atoms of iron unite and form a double atom which then forms compounds with other elements. A moment's reflection will show that this double atom, formed from two atoms each having a capacity of four, will have a power of six, one degree of atomicity in each atom having been consumed in forming the compound.

For all these cases of varying atomicity, whether regular or irregular, the terminations *ous* and *ic* are much employed, *ous* indicating the lower degree and *ic* the higher. We have, in this way mercurous (lower atomicity) and mercuric (higher atomicity) salts; ferrous (dyad) and ferric (hexad) compounds. Indeed, in the use of the termination of the acids, the same principle is carried out, sulphurous acid being the compound in which sulphur has a lower (tetrad) atomicity; sulphuric acid, one in which sulphur has a higher (hexad) power.

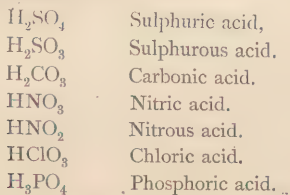
In arranging formulæ containing three elements, the application of the law of atomicity becomes somewhat difficult. In many common cases, it will be found that the atomicity of one of the elements is much higher than it is in bodies containing two elements; sulphur, for instance, is a dyad in binary compounds, but in the sulphites and sulphates it is, respectively, a tetrad and hexad. When oxygen is one of the three elements we usually count it against the sum of the other two. Taking an instance of the salts above mentioned, we would get the following formulæ :



In each case the atomicity of the oxygen is equal to the sum of that of the other two elements.

For the student, however, the safest and shortest rule will be to commit to memory, thoroughly, some standard formulæ containing three elements, and from these, by very simple rules, a large number of compounds can be built up.

These formulæ are :



The derivatives from these bodies form a large part of common chemical substances. If we wish to write the formula of any

metallic salt, we substitute the proper amount of metal for the hydrogen in the corresponding acid. Let it be required to write the formula of potassium carbonate; the reasoning would be as follows: Carbonic acid is H_2CO_3 , potassium is a monad; two atoms of potassium will be required to substitute the two atoms of hydrogen, and the formula is K_2CO_3 . By the same reasoning, copper sulphate may be deduced. Sulphuric acid is H_2SO_4 , copper is a dyad; one atom of copper will displace two of hydrogen, therefore, CuSO_4 . When the standard formula contains too small an amount of hydrogen, we must multiply the expression by some whole number. For instance, the formula of copper nitrate will be deduced in this manner: Nitric acid is HNO_3 , copper is a dyad; copper will therefore replace the hydrogen of two molecules of nitric acid; hence, $\text{Cu}(\text{NO}_3)_2$ or CuN_2O_6 .

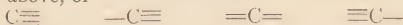
Some chemists, in order to avoid this latter difficulty, use always the double formulæ and write as follows:

Nitric acid	$\text{H}_2\text{N}_2\text{O}_6$
Nitrous acid	$\text{H}_2\text{N}_2\text{O}_4$
Chloric acid	$\text{H}_2\text{Cl}_2\text{O}_6$
Phosphoric acid	$\text{H}_6\text{P}_2\text{O}_8$

Graphic Formulæ. A convenient and much used method of indicating atomicities is by GRAPHIC FORMULÆ. These consist of the symbol of each element with bonds or prolongations, the same in number as the degrees of atomicity. Taking some common elements as examples, we have

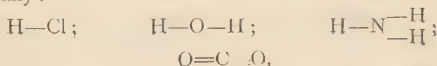
monad	dyad	triad	tetrad	pentad
$\text{K}—$	$—\text{O}—$	$—\text{P}—$	$—\text{C}—$	$=\text{N}=$

These bonds may be attached in any position or direction as long as the proper number is used. Carbon, for instance, may be written as above, or



or in any other way, provided four bonds are present.

The practical application of this graphic notation to the writing of chemical formulæ is easy. We link together the bonds of the different elements, and when all the points are joined, the compound is complete, and is a saturated molecule. Two bonds of one atom, however, can never be attached to a single bond of another atom. We cannot have $\text{K}=\text{O}$, but $\text{K}—\text{O}—\text{K}$ showing us that the composition of potassium monoxide must be K_2O . The following are examples of some common compounds written graphically:



We may, also, indicate *unsaturated* molecules. Thus, $O=C=$ shows that carbon monoxide is a body having two degrees of atomicity unsatisfied ; $O=C=\overset{Cl}{\underset{Cl}{| \quad |}}$ that two atoms of chlorine have combined and satisfied this free atomicity.

The only objection, perhaps, to the use of graphic formulæ is the danger that the student may think that the atoms actually have spokes, or projections on them, or are arranged in the somewhat architectural manner shown in the formulæ. $K-O-H$ does not mean that in potassium hydrate the atom of O is flanked on either side by a potassium, and by a hydrogen atom, or that the atoms are connected by hooks or prongs ; but merely that the oxygen atom has certain degrees of affinity, which are satisfied by other atoms. A special application of this notation is to explain the nature of those changes in atomicity which have already been mentioned. The atom of sulphur, for instance, is, in some combinations, a hexad, in others a tetrad, in others a dyad. This progressive diminution of capacity may be supposed to arise from the bonds of affinity combining with each other in pairs, thus :

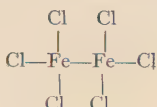
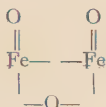


The same principle can be shown with an element of uneven atomicity.



Since such combinations cannot take place unless both points are saturated or neutralized, the decrease of atomicity must take place by two degrees.

The nature of the change by which the iron atom passes from a dyad to a hexad condition, can be very well shown by this method. Dyad iron, graphically represented, would be $(Fe=$, in which two bonds have satisfied each other, leaving two still active. In the higher degree of atomicity, the condition is $-Fe-Fe-$, one bond of each atom having combined and linked the two in chemical union. Ferric oxide and ferric chloride would be



6. ELECTRICAL RELATIONS OF THE ELEMENTS.

ELECTRICAL excitement exhibits two opposite conditions, called respectively, **POSITIVE** and **NEGATIVE**. These conditions are produced in any apparatus developing electricity. The points at which the electrical excitement is manifested, for instance, the wires of a battery, are called the **POLES**. The positive pole is usually distinguished by the sign $+$, and the negative by the sign $-$.

These have a strong tendency to unite and neutralize each other. On the other hand, positive electricity repels positive, and negative repels negative. Two bodies charged with different kinds of electricity will attract each other, but if charged with the same kind of electricity will repel. The law is generally expressed as follows: *like electricities, repel; unlike, attract.*

These principles have been applied to the determination of some important relations between elements. A current of electricity decomposes a large number of compound bodies, and certain elements are strongly attracted to the positive pole, and others to the negative. Thus, potassium will always be liberated in contact with the surface negatively charged, while oxygen will be set free in contact with the positive surface. This will be the invariable result with these elements, no matter what compounds be taken for the experiment, but with many other elements the effect will depend upon the nature of the compound. With H_2S the sulphur will be attracted by the positive pole; with SO_2 by the negative. This difference is due to the superior attraction which the oxygen has for the positive pole, by which it seems to compel the sulphur to go to the other point.

Since unlike electricities attract, it follows that elements which are drawn to the positive side must be negative, and those drawn to the negative side must be positive. Very frequently we use the syllable "**ELECTRO**" in this connection,—thus we say zinc is *electro-positive*; chlorine is *electro negative*.

A body is not absolutely positive, or absolutely negative, but is simply more positive or more negative than some other substance. Nevertheless, as the list of elements is limited, we will have two bodies, which, by their high affinities, will stand at the extremes of the scale, one being always negative, the other always positive. Leaving out of consideration some rare elements, we may place potassium as the most positive, oxygen as the most negative.

The following table shows the common elements arranged in the order of their electrical relations:

K	Fe	Hg	S
Na	Ni	H	N
Ba	Pb	Sb	I
Ca	Sn	C	Br
Mg	Cu	As	Cl
Zn	Ag	P	O

It is not necessary to commit this table to memory. The general purpose is met by remembering that the metals are positive, and the non-metallic elements negative.

Each element of this list will be positive when in combination with any element below it, negative when in combination with any above it.

These principles have been explained at some length, because they find an important application in determining chemical changes. The greatest chemical attraction exists between elements most widely separated in their electrical relations. The student may, by this means, often discover the most probable result of any chemical action, as will be fully shown in another section.

7. REACTIONS.

CHEMICAL symbols are employed not only to show the composition of bodies, but also to show exactly the nature of the chemical changes which occur when different bodies are brought in contact. When so used, the expression is called a REACTION. Certain compounds, which are much used for producing reactions, are called REAGENTS, though strictly all the substances present take equal part in a reaction. When we pour vinegar upon a marble table, we say, in ordinary phrase, that the marble is corroded, but, in fact, the vinegar is equally acted upon, both substances are changed in composition, both are rendered unfit for their original uses; in other words, they have not only acted, they have reacted, and are therefore both reagents.

A reaction is substantially an expression of the results of an experiment, and, when correctly written, gives us the proportion in which bodies are to be used and the proportion of the resulting substances. Speaking absolutely, we can never be sure of the correctness of any reaction until we make the experiment and analyze the result; but the progress of chemistry has made known certain laws of change, which enable us to predict, or infer, many results without the necessity of actual observation. Every now and then, however, the analogy fails, and experiment disappoints the suggestions of theory.

Reactions are written by placing in proper proportion and connected by \div signs the formulæ of the bodies concerned, then writing the sign $=$ and following this by the formulæ of the resulting bodies. For instance,



expresses that on bringing together silver nitrate and hydrochloric acid, a chemical change occurs by which silver chloride and nitric acid are produced. Students find, in regard to writing reactions, three difficulties: 1st. To know whether a given change will take place. 2nd. To know the quantities of the bodies to be used. 3rd. To know the nature of the resulting bodies. These difficulties may be taken up in order. 1st. In the simplest cases, the nature of the reaction will be determined by the affinities of the elements as governed by their electrical relations, the change taking place in such a way that the element having the stronger electric affinity will drive out and supplant the element of similar but weaker affinity. A reference to the table of electrical affinities will show that chlorine is more strongly negative than bromine, and bromine than iodine. Accordingly, we find that when chlorine acts upon the bromides they are decomposed, the bromine being expelled, and that bromine, in turn, expels iodine from combination. Therefore, such reactions as



are simply illustrations of the general electrical relations of elements concerned. If these affinities were the only active causes of chemical change, the subject would be quite simple, but, by repeated experiment, chemists now know that the surrounding circumstances may suspend or modify the play of affinities, so as to produce an endless variety of chemical action. It is hardly necessary to remark that all the modifying influences are not yet known. Three of them, however, are pretty well understood, and are of such importance as to require the attention of the student.

(a) **Insolubility.** When in any liquid we bring together substances which are capable of forming a body insoluble in the liquid, that insoluble compound will be produced in spite of the general relations of affinities. This influence of insolubility is the basis of a large number of tests and other chemical operations.

When the formation of the insoluble compound would require a powerful chemical agent to be set free, the change will not take place, unless, of course, the added substance is stronger than the one to be liberated. Carbonic acid forms with calcium a body quite insoluble in water, but this body cannot be formed by passing carbonic acid into calcium sulphate. The reason is shown

at once on examining the conditions of the experiment. The reaction would have to be $\text{CaSO}_4 + \text{H}_2\text{CO}_3 = \text{CaCO}_3 \text{ (insoluble)} + \text{H}_2\text{SO}_4$, that is sulphuric acid would be set free. The affinity of H_2CO_3 is, under ordinary conditions, so much below that of H_2SO_4 that the former will not drive out the latter. The condition becomes changed if we assist the action of the carbonic acid by some substance which has an affinity for sulphuric acid and will prevent it being set free. $\text{CaSO}_4 + \text{Na}_2\text{CO}_3$ will produce immediate action, resulting in $\text{CaCO}_3 + \text{Na}_2\text{SO}_4$. This reaction illustrates a common method of keeping the powerful affinities in abeyance, and thus allowing secondary influences full play. Some of the arsenic tests show the principle strikingly. Arsenious acid added to CuSO_4 produces no action because the affinity of the SO_4 is too strong, but by adding a little ammonia, the strong affinity this has for SO_4 assists in breaking up the copper sulphate and immediately a precipitate of copper arsenite falls.

(b) **Volatility.** This is the second influence that disturbs ordinary affinities. If a body is capable of being converted into a gas, this fact will diminish its chemical power; fixed substances that have ordinarily less affinity will drive it out of combination. Boric acid, for instance, is one of the weak acids, yet at a red heat it will drive out even sulphuric acid. The cause is, in the main, that at this temperature sulphuric acid is volatile, while boric acid is fixed. Chemists make, as is well known, much use of the action of heat as a modifier of chemical action, and frequently it is this influence of volatility which is brought into play.

The ease with which hydrogen is driven out of combination may be regarded as due to its volatility, it being a gas even at low temperatures.

(c) **Mass.** Sometimes chemical action seems to be governed by the quantity of the substance present. If we pass water vapor over red hot iron, iron oxide is formed and hydrogen is set free; if we pass the hydrogen back over the iron oxide, steam is formed and iron set free. In the first case, the water is in excess and exerts an oxidizing influence; in the second, the hydrogen is in excess, and exerts a deoxidizing influence. The effect of mass is indefinite and uncertain, and need not enter into the ordinary working of reactions.

It will be seen to be a deduction from these statements that no substance can be set down as absolutely the strongest in affinity. Chemists cannot determine, for instance, what is the strongest acid or the strongest alkali, except under specified conditions.

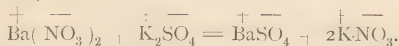
2nd. The proportion in which bodies react is determined by their atomicities. Let it be required to write the reaction between mercuric chloride and potassium iodide. The formulæ

are HgCl_2 and KI , but the bodies will not react in this proportion for the Hg will require I_2 and Cl_2 will require K_2 . The proper reaction is $\text{HgCl}_2 + 2\text{KI} = \text{HgI}_2 + 2\text{KCl}$. In the same way, antimony sulphide and hydrochloric acid can only act upon each other in the ratio $\text{Sb}_2\text{S}_3 + 6\text{HCl}$ because Sb being a triad, Sb_2 will combine with Cl_6 , and S being a dyad, S_3 will require H_6 .

3rd. If a chemical change occurs when two given substances are brought in contact, the nature of it will depend principally upon the electrical relations of the elements concerned. In the reaction $\text{HgCl}_2 + \text{H}_2\text{S}$, the only possible result is the combination of S with Hg and H with Cl , as is shown at once by placing

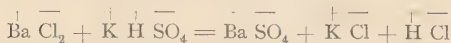
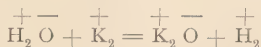
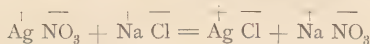
the proper signs over the elements, $\overset{+}{\text{Hg}} \overset{-}{\text{Cl}}_2 + \overset{+}{\text{H}}_2 \overset{-}{\text{S}}$. Such a

combination as $\overset{+}{\text{Hg}} \overset{+}{\text{H}}_2$ or $\overset{-}{\text{Cl}}_2 \overset{-}{\text{S}}$ could not take place, since it requires like electricities to attract, which is against the rule. In beginning with reactions, the student will do well to place the proper signs over each element, and these signs will be a useful guide and control. When acids or salts, containing three elements, are part of the reaction, the plus sign is put over the hydrogen or the metal, and the negative sign over the oxygen and other element, thus :



The placing of the single sign over the two elements is simply an evidence of the fact that in ordinary reactions, these two elements act as a single element.

The following formulæ will further illustrate the general principle :



In the last reaction, the electro-positives K and H may seem to be in union, but this is not the case. Each is independently united to the SO_4 , which is a dyad. The formula might be

written $\overset{+}{\text{K}}\overset{+}{\text{H}}\overset{-}{\text{SO}}_4$

In writing reactions in which any element in the free state is expressed it is now customary to use such proportions as will give an even number of atoms of the element. Thus, in giving the reaction of sodium upon hydrochloric acid, it would be written



and not



This system has been adopted in deference to the theory that no atom exists alone, but that elementary bodies have their atoms united in pairs. No advantage of any practical character accrues to the student from this complication, and in the present work the simpler method has been adopted. It will occasion no inconvenience, because, if the reaction is known and understood in the one form, an easy multiplication will change it into the other.

8. COMBINATION BY VOLUME.

IN THE section on the laws of chemical combination, mention was made of the fact that the elements combined with each other in definite proportions; generally in the proportion of the atomic weights, or some multiple of them. Modern chemical research has called attention to important facts in regard to the proportion by *volume* or *bulk* in which the combination takes place. As yet, no very exact or important results have been obtained from the study of the volume in the solid or liquid form, but in the state of gas or vapor, the relations of the different elements are quite striking.

If we weigh equal volumes of the elements in the state of gas, we find that their relative weights will, with a few exceptions, be in exact proportion to their atomic weight. For instance, a vessel which holds 1 grain of hydrogen (about 47 cubic inches) will hold the following quantities of other elements, it being understood that all the bodies are in the state of gas, and at the same temperature and pressure:

Element.	Atomic Weight.	Wt. of vol. equal to 1 vol. of H.
O	16	16
S	32	32
Cl	35.4	35.4
I	127	127
Br	80	80

Some of the elements cannot be converted into vapor and consequently cannot be compared on this system. Among these are

carbon, silicon, and many metals. These practically resist the action of the highest temperature which can be used in such experiments. A few elements show results which are exceptional. Among these are:

Element.	Atomic Weight.	Wt. of vol. equal to 1 vol. of H.
As	75	150
P	31	62
Hg	200	100

In the case of phosphorus and arsenic the weight is twice as great as what analogy would require; in the case of mercury, half as great.

Changes in temperature and pressure, produce changes in the volume of gases, and all gases are affected nearly equally. It was pointed out, on page 2, that a change of volume is due to a separation of the molecules or atoms, and not to any change in the volume of the atoms themselves. It is a reasonable assumption that, if two gases expand equally under the same conditions, it is because they contain the same number of inter-spaces in which the expansion takes place. Hence a law which is expressed thus: Equal volumes of elementary gases contain equal numbers of molecules.

The weight of the atoms of each element may be determined by this law. If a given volume of hydrogen contains, say 1,000 molecules, the same volume of oxygen will contain the same number, and as the oxygen volume is 16 times as heavy as the hydrogen, it is clear that the weight of each molecule of oxygen will be 16 times that of each molecule of hydrogen. The molecules of hydrogen and oxygen each contain two atoms, hence the atomic weights will also be in the proportion of 16 to 1.

As gases are decidedly affected by even slight changes in temperature and pressure, it becomes necessary to refer all observations to a standard condition. Under the English system of weights and measures this has usually been 60° F. for temperature, and 30 inches of mercury, about 15 lbs per square inch, for pressure. Under the French system the temperature is 0°C (sometimes 15°C) and 760 millimetres of mercury. This latter figure is practically the same as 30 inches, and is the height of the barometer at the level of the sea.

A study of the condition of the ordinary gases leads us to believe that the spaces between the molecules are very much greater than the molecules themselves. Ample room exists in any gas for adding atoms or molecules without increasing the volume. The phenomena of combination between gases seem to show that all molecules are of the same size; at least it is known

that the elementary gases combine in such a way as to produce a volume of gas *which is equal to twice the volume that would be occupied by one atomic weight of hydrogen*. The following instances are taken from among the commonest chemical compounds:

One volume of H and one volume of Cl combine and produce *two* volumes of HCl.

Two volumes of H and one volume of O combine and produce *two* volumes of H_2O .

Three volumes of H and one volume of N combine and produce *two* volumes of NH_3 .

In each of these cases it is seen that the volume of the resulting combination is twice that of the one atomic weight of hydrogen. If the weights should be in grammes, then the resulting volume would be that occupied by two grammes of H; if the weights should be in grains, then the resulting volumes would be that occupied by two grains of H; and if the substances were estimated by volume alone, say in pints, then the resulting compounds would have the volume of two pints.

The great majority of chemical compounds that have been examined conform to this law. A few exceptions are known, and as mentioned above, some elements cannot be converted into gas and thus cannot be examined on this point. It follows, from this uniform condensation, that when the molecule contains many atoms the gas will be heavy, and further, that its weight compared to hydrogen can be easily calculated.

Some examples will make this plain:

47 cubic inches of H, weighing 1 grain, will combine with 47 cubic inches of Cl, weighing 35.4 grains, and produce 94 cubic inches (*i. e.* 47×2) of hydrochloric acid (HCl) weighing 36.4 grains, and by dividing this last result by 2, we get the weight of a quantity of hydrochloric acid equal to the one atomic weight of hydrogen, viz. 18.2. This figure, 18.2, represents, therefore, the density or specific gravity compared to hydrogen.

94 cubic inches of H, weighing 2 grains, will combine with 47 cubic inches of O, weighing 16 grains, and produce 94 cubic inches of steam, H_2O , weighing 18 grains. If we divide 18 by 2, we get, as before, the density of steam compared to hydrogen, viz. 9.

47 cubic inches of N, weighing 14 grains, will combine with 141 cubic inches (47×3) of H, weighing 3 grains, and form 94 cubic inches of ammonia, NH_3 , weighing 17 grains, and this weight divided by 2 gives 8.5 as the density of ammonia compared to hydrogen.

The instances might be multiplied to a great extent, but it is sufficient for the student to know that the facts have been well established, and that very few exceptions to the rule are now known.

It has also been found that the amounts of heat required to raise equal weights of the different elements to the same temperature bears a definite relation to the atomic weight, and may be used as a means of determining it. Hydrogen requires sixteen times as much heat as oxygen, when both are under precisely the same conditions.

DESCRIPTIVE CHEMISTRY.

IN the following descriptive part the arrangement of the elements has not been made in strict accordance with accepted classifications. It has been deemed best to describe at the outset the more common ones, in order that the chemistry of the more important compounds may be given early in the work.

HYDROGEN, H.

ATOMIC WEIGHT I. DENSITY I.

Sources. Exists in water, and in all animal and vegetable substances. It was discovered by Cavendish, in 1776. The name means "producer of water."

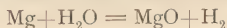
Preparation. 1. Directly from water by the action of various metals. Sodium and potassium act in the cold; iron, magnesium, zinc, etc. require a high temperature. Carbon may also be used, and is especially suitable when very large quantities are required.

2. By a sort of galvanic action of alkaline solutions on a mixture of zinc and iron filings, or on aluminium.

3. From most acids, by the action of common metals, generally without heat.

Exp.—A piece of sodium amalgam (q. v.) is placed under an inverted test tube filled with water and standing on the pneumatic trough. The gas collects in the tube and may be tested as explained below. The water becomes alkaline from the formation of caustic soda. $\text{Na} + \text{H}_2\text{O} = \text{NaHO} + \text{H}$.

Exp. Magnesium ribbon is folded once or twice, and placed in a small, hard-glass tube, one end of which is partially drawn down to a jet, the other attached to an apparatus for producing steam. Steam is passed through until the tube and ribbon are free from condensed moisture. The metal is then strongly heated at the extreme point; it takes fire, producing an oxide and free hydrogen, which latter can be burned at the jet.



Exp. Fragments of zinc are introduced into the gas bottle and dilute sulphuric or hydrochloric acid added; the gas comes off freely, and being very light, and insoluble in water, may be collected either by upward displacement or by the pneumatic trough.



Properties. Pure hydrogen is colorless, tasteless and inodorous. It is the lightest body known, a litre weighing 0.08961 gm. It can be liquified only by intense cold and pressure.

A mixture of hydrogen with air explodes when lighted, hence no experiment should be made until all the air is driven out of the evolution bottle. The best test is to collect a small quantity in an inverted test tube, by displacement. After the gas has passed in for about half a minute, the tube should be removed, and still keeping it inverted, a light should be applied to its

mouth; a sharp explosion, extending at once through the tube, indicates the presence of air; a slight explosion at the mouth of the tube only, and a faint flame moving gradually up the tube, shows the gas to be pure.

The important properties of hydrogen may be shown by the following experiments :

1. Lighted at the mouth of the delivery tube it burns with a pale blue, but very hot flame. If the jet is of glass, the flame becomes yellow; a dry vessel held over it becomes coated with moisture, which is the sole product of the combustion.

2. A stout, wide-mouth bottle, or better, a small, tin vessel filled with a mixture of gas and air, gives a loud explosion on being ignited.

3. Soap bubbles blown with the gas, rise rapidly in the air.

4. A large, light vessel being counterpoised on a delicate balance, will be thrown decidedly out of balance by substituting hydrogen for the contained air, by displacement; on placing the vessel with the mouth upward, the hydrogen quickly escapes, and the equilibrium is restored.

Hydrogen, though not poisonous, will not sustain life; small quantities may be inhaled without danger, and produce a peculiar change in the voice; for this the gas must be absolutely pure, which that made from commercial articles never is.

Gen. Chem. Rel. Hydrogen is electro-positive and is a standard for atomicity, atomic and molecular weight and density. It combines with many elements. It seems to be capable of forming alloys with some other metals, especially palladium and platinum. As explained in the section on nomenclature, hydrogen is regarded as the essential element of all acids. Keeping in view the fact that the affinity of substances is diminished by volatility, it can be easily understood that the hydrogen compounds should be easily decomposable and should be active chemical agents. Such a body as H_2SO_4 is the most active of the sulphates, because its positive element (hydrogen) is of low affinity, and is displaced by a large number of bodies.

OXYGEN. O.

ATOMIC WEIGHT 16. DENSITY 16.

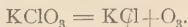
Sources. Exists in water, air, all animal and vegetable tissues, and in the great majority of minerals. Constitutes over half the matter composing the earth. Discovered by Scheele and Priestley in 1774. The name means "producer of acids."

Preparation. 1. The oxides of mercury and of some other metals are decomposed by heating. This method is of little practical value, but is interesting as it was the means of the discovery of the gas. $\text{HgO} = \text{Hg} + \text{O}$.

2. Certain compounds of manganese and barium when heated in a current of air, absorb oxygen, and give it out in a current of steam. By alternating these two currents, large quantities of oxygen may be obtained. Such processes are not suitable for laboratory work.

3. The chlorates, nitrates and some rarer salts are decomposed by heat, giving off large quantities of oxygen, but not always quite pure. Potassium chlorate is by far the most suitable. Used alone it requires a high temperature, but when mixed with about one-quarter of its weight of manganese dioxide, a heat of about 500°F . (260°C) is sufficient. The exact manner in which manganese dioxide acts has not been explained.

Exp. Mix thoroughly 4 parts potassium chlorate with 1 part manganese dioxide, and heat the mixture in any suitable vessel. One ounce of potassium chlorate yields nearly two gallons of the gas. It may be collected by downward displacement, but is not quite pure; it is best collected over the pneumatic trough.



Properties. Pure oxygen is colorless, odorless and tasteless; it is one-tenth heavier than air, one litre weighing 1.43 gm. It is continually being absorbed by living animals, in the process of respiration, to which function it is essential. It is also consumed in all ordinary combustion. The pure gas causes considerable excitement of the vital functions of animals, and greatly increased action in ordinary flames, but does not, as a rule, produce spontaneous combustion.

Exp. A taper, or splinter of wood, lighted, and blown out in such a way as to leave a glowing coal, is instantly relighted, with a slight explosion, on putting it into the gas. The experiment may be repeated several times.

Exp. A bit of bark, or knot of charcoal ignited at one point and plunged into the gas, burns brilliantly, producing a colorless gas which combines with water and forms an acid.

Exp. Sulphur burns in oxygen with moderate brilliancy, producing a colorless, highly irritating gas, which in combination with water produces an acid. A picture painted with a solution of quiniæ sulphate is almost invisible by ordinary light, but becomes visible when illuminated by burning sulphur; the explanation of this effect belongs to physics; it is called fluorescence, and is shown by many substances.

Exp. Phosphorus burns in the gas, with great brilliancy, producing dense white clouds which rapidly absorb water and produce a powerful acid.

It will be noticed that in the foregoing experiments acids have been produced by the combustions. Lavoisier supposed that oxygen was necessary to the production of an acid; the present name of the gas means "acid producer," and expresses this view. According to the opinion of modern chemists, hydrogen is the essential element of acids, and the following experiments will show that the combination of oxygen may produce bodies very different from acids.

Exp. Magnesium ribbon ignited and plunged into oxygen, burns rapidly and brightly, producing a bulky, white powder, insoluble in water and of alkaline properties.

Exp. One end of a thin, steel ribbon (watch spring) is wrapped with a few turns of cotton thread and then dipped into melted sulphur or wax. This end being lighted, the ribbon is put into oxygen; the metal quickly takes fire, produces an abundance of sparks, and from time to time drops hot globules to the bottom of the jar, which often fuse themselves into the glass. A layer of three or four inches of water or an inch of sand will generally prevent the breaking of the jar. The product is black, insoluble in water, and destitute of acid properties.

Sodium, potassium and zinc turnings also burn in oxygen with more or less brilliancy.

The explanation of the production of acids in the first set of experiments is that the water, which is necessarily present, forms with the product of combustion a new substance. The combinations of oxygen are called oxides; in the case of the burning sulphur, the resulting gas has the formula SO_2 , sulphur dioxide, and if the materials were perfectly dry, no acid would be formed. In the presence of water we have the reaction $\text{H}_2\text{O} + \text{SO}_2 = \text{H}_2\text{SO}_3$, sulphurous acid. Similarly the charcoal produces CO_2 which uniting with water produces H_2CO_3 , carbonic acid. The same principle applies to the phosphorus experiment.

On the other hand, the metals, manganese, sodium, etc., form oxides which so far from being acids, are really powerful neutralizers of acids, and are said to be alkaline or basic. They combine with water, producing compounds which retain the alkaline character. In this way we have $\text{MgO} + \text{H}_2\text{O} = \text{MgH}_2\text{O}_2$, $\text{Na}_2\text{O} + \text{H}_2\text{O} = \text{Na}_2\text{H}_2\text{O}_2$. Further explanations will be found in section on water.

Gen. Chem. Rel. Oxygen combines with every other element except fluorine, and with many, in several proportions. The chemical relations of these oxides are dependent in part up-

on the number of oxygen atoms present. With the non-metals this effect is not well seen, but a number of metallic compounds show it distinctly. The oxides of manganese may be taken as examples :

MnO	Powerful base.
Mn_2O_3	Weak base.
MnO_2	Indifferent.
MnO_3	Forming an acid.

These instances illustrate the general law that small proportions of oxygen tend to produce bases, high proportions anhydrides or acid-forming oxides, and intermediate proportions, bodies of uncertain or indifferent character. The non-metallic elements being apparently incapable of yielding bases, form with oxygen, in low proportions, neutral oxides. This is shown by the nitrogen series :

N_2O	Indifferent.
NO	Indifferent.
N_2O_3	Acid-forming.
NO_2	Doubtful.
N_2O_5	Acid-forming.

Oxygen is a dyad, and is generally considered the most electro-negative element, but under some circumstances chlorine seems to be superior in this respect. It is slightly soluble in water, and upon this fact depends the existence of most forms of aquatic life.

Substances which take away oxygen from its combinations are called *reducing agents* ; those which add oxygen, *oxidizing agents*.

Tests. Besides the power of relighting a taper, free oxygen may be recognized by turning brown a mixture of caustic soda and pyrogallin, and by converting colorless nitric oxide into red peroxide.

Ozone. Oxygen is susceptible of a modification of some of its properties without chemical change. The study of this subject was begun in 1840 by Schoenbein, and very many researches have since been made without, however, completely explaining the condition. The modified oxygen is called 'ozone' from a Greek word meaning "to smell," on account of its marked odor.

Ozone may be prepared in several ways :

1st. By a succession of electrical sparks through air or oxygen. Its peculiar odor is observed when an electrical machine is put into active operation.

2nd. By the slow oxidation of phosphorus, and of turpentine and other essential oils.

3rd. By the decomposition of water by the galvanic current.

4th. By the action of acids upon certain bodies rich in oxygen.

By all these methods, only a small proportion of the oxygen is converted into ozone.

Exp. Place a few crystals of potassium permanganate in a wide mouth bottle, and add a few drops of sulphuric acid. A peculiar odor is noticed; the evolved gas will tarnish mercury and silver, and turn blue a piece of paper soaked in a solution of potassium iodide and starch. This last effect is due to the setting free of the iodine by the superior affinity of the ozone for the potassium. Organic matter is also acted upon powerfully.

Ozone is heavier than oxygen, is soluble in water, is converted into common oxygen by heat and by some metallic oxides. It is generally present in the atmosphere, especially in open country places. It is considered an important natural disinfectant.

Another modification of oxygen called antozone has been supposed to exist, but this is considered by many chemists to be hydrogen dioxide, H_2O_2 .

The Compounds of Hydrogen and Oxygen are two:—

H_2O	Water.
H_2O_2	Hydrogen Dioxide.

Water is extensively distributed in nature. Besides being collected into large masses, oceans, lakes, etc., it exists in suspension in the atmosphere, in most minerals, and in all animal and vegetable tissues. Some living structures, such as succulent fruit, jelly fish, etc., consist almost entirely of water.

Water rarely occurs pure in nature. Rain contains traces of ammonium salts; rivers and springs are contaminated generally with sulphates, chlorides, carbonates and organic substances. The total quantity ordinarily dissolved varies from five to thirty grains to the gallon. When the quantity greatly exceeds this, and especially when peculiar substances, such as iron or sulphur, are present, it constitutes a *mineral water*. Sea water is quite rich in mineral substances, and may be regarded as a mineral water.

The more important varieties of natural waters are:

Alkaline or Carbonated waters, containing various carbonates in solution, generally with a quantity of free carbonic acid.

Hard waters in which a decided amount of magnesium or

calcium salts is present. The chemical relations of these waters will be hereafter explained.

Chalybeate waters, containing iron, generally as ferrous carbonate with excess of carbonic acid.

Sulphur waters, containing sulphureted hydrogen and other sulphides.

Acid waters, containing some of the stronger acids in the free state.

Saline or Aperient waters, having large amounts of chlorides and sulphates.

The substances thus dissolved in water have influence upon the health of those regularly drinking it, but as to the exact nature of these effects little is known. Springs of chalybeate, alkaline and other mineral waters are resorted to by invalids, but how much of the apparent beneficial action is due to the water and how much to change of scene, diet and other conditions, is a mooted point. With regard to ordinary drinking water, its composition may vary within wide limits without appreciable effect upon those who drink it. According to the generally accepted method, the important points to be determined in the analysis of water are: total amount of dissolved matter, degree of hardness, amount of chlorine, nitrates and nitrites, and of organic matter. The chlorine and organic matter are believed to represent generally the amount of refuse matter or sewage, the nitrates and nitrites, the "previous sewage contamination" or organic matter, which has been altered by oxidation. The most difficult problem is the correct determination of the organic matter. By this term is meant the animal and vegetable substances that get in with the waste, drainage and sewage which all flowing water receives. These substances are in minute quantities, even in impure water, but appear to be capable of doing much harm. Three principal methods are now used for their estimation:—

(a) Frankland's, in which the amount of carbon and nitrogen is determined.

(b) Wanklyn's, in which the nitrogen is converted into ammonia and estimated.

(c) Tidy's modification of Forchammer's, in which the amount of oxygen required to oxidize the organic matter is determined. The details and discussion of the value of these methods belong to analytical chemistry.

The originator of each process has stoutly defended his own, and strongly condemned the other methods,

Purification of water. Filtration through porous materials,

paper, charcoal, sand, etc., will remove only suspended impurities. Animal, charcoal and spongy iron remove some of the dissolved matters, but distillation is the only method of preparing absolutely pure water.

Exp. Boil some ordinary water with fragments of dried leaves or hay, add some common salt and fine chalk, and shake. The specimen now represents a highly impure and turbid water; filtered through paper it becomes clear, but it still contains the salt and the organic matter. Filtered through animal charcoal, (q.v.), it will lose the organic matter, at least to some extent. The difference may be shown by adding to samples of the water before and after filtration a few drops of a solution of potassium permanganate. The sample which has not passed through charcoal will rapidly decolorize permanganate; the filtered sample will have less action. In the first case the organic matter abstracts oxygen from the permanganate and decomposes it. Put the filtered liquid into a retort, provided with a condenser, and boil until a perceptible quantity of liquid collects in the receiver; it is almost chemically pure, containing only a minute quantity of ammonia.

Quantitative composition of water. When two volumes of hydrogen and one volume of oxygen are combined, complete condensation takes place, and water is formed. When water is decomposed by the electrical current, two volumes of hydrogen and one volume of oxygen are always obtained. Since oxygen is sixteen times as heavy as hydrogen, the proportion by weight will be 2 to 16 or 1 to 8. The formula H_2O indicates that the molecule contains two atoms of hydrogen and one atom of oxygen. When the combination of the two gases is performed at a temperature of $212^{\circ} F.$ ($100^{\circ} C.$) or over, the water remains as steam, equal in volume to the original hydrogen. The theoretical relations of this fact have been considered in the section on combination by volume. The composition of water *by weight* has been established with great accuracy by passing hydrogen over hot copper oxide; the oxygen is abstracted from the copper and combines with the hydrogen. The water so formed is collected and weighed; the loss of weight of the copper oxide gives the quantity of oxygen taken up, and the difference is the hydrogen. The actual result of such an experiment, conducted with every precaution to insure accuracy, is:

Hydrogen	11.11
Oxygen	88.89
	<hr/>
	100.00

which is substantially 1 to 8.

Properties and Chem. Rel.—Pure water is odorless and tasteless, and in large masses distinctly blue. Under ordinary pressure it freezes at 32° F. (0° C.), and boils at 212° F. (100° C.). It dissolves many bodies, solid, liquid and gaseous, and is absorbed by many substances. Porous bodies, for instance, exposed to moist air increase in weight considerably by absorbing water. These effects are not generally regarded as due to chemical affinity. Solution of solids in water is generally attended with the production of cold, as may be shown by making a solution of Epsom salt or ammonium chloride. The solution of gases is apt to be attended with the production of heat. The solvent power of water is quite extensive, though very different for different bodies. As a rule, the solubility of solids is increased by heat; that of gases diminished. Gases exhibit great differences; some *e. g.* hydrogen and oxygen, dissolve very sparingly; ammonia and hydrochloric acid, on the other hand, are largely dissolved. Gases may all be expelled by thorough boiling. When strong solutions are prepared by heat, the liquid on cooling, usually deposits a portion of the dissolved substance in a crystalline form.

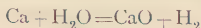
Exp. Bring some water to boiling, and introduce, by small portions at a time, potassium chlorate, until no more dissolves; pour the solution into a thin glass beaker, and the chlorate will be continually deposited in fine scales until the liquid is quite cold. Lead iodide treated in the same way gives even a better effect.

A very large number of substances form, with water, definite chemical compounds, in some of which the water unites directly without losing its identity; in others the molecule H_2O seems to be broken up. Of the first kind of combination, the best instances are in the crystallization of the common metallic salts. Copper sulphate, for instance, has the formula CuSO_4 , but the blue crystals sold in the shops as copper sulphate contain, in addition to this formula, a large amount of water. This water is derived from that in which the copper sulphate was dissolved. It is an essential part of the crystal, for, if we drive out all of it, the mass is converted into a white powder. Chemical analysis shows that the composition of the blue substance is $\text{CuSO}_4 + 5\text{H}_2\text{O}$. Water that is in this way part of a molecule and essential to a crystalline form is called WATER OF CRYSTALLIZATION. Substances that do not contain it in such a state of combination, are said to be ANHYDROUS. Some metallic salts form crystals containing large amounts of water; sodium carbonate in its commercial form, $\text{Na}_2\text{CO}_3 + 10\text{H}_2\text{O}$, contains over fifty per cent. water.

The second form of the chemical action of water is where it

is apparently decomposed, its two elements associating themselves independently with the elements of the other body. If we mix water with common quick-lime, a violent action, attended with production of much heat, ensues, and if the quantity of water is about one-third that of the lime, a dry powder results, from which no appreciable amount of water can be expelled except by a red heat. If this compound contained water in an unchanged form, a moderate heat would drive it all out; hence, chemists have regarded the compound not as CaO , H_2O but as CaH_2O_2 which is an entirely new body, containing really neither quick-lime nor water. A considerable number of oxides are capable of uniting thus with water and forming bodies known as **HYDRATES**.

Perhaps the most scientific view of these hydrates, is to regard the water as acting the part of an acid (it might indeed be called hydric acid), and just as CaSO_4 is calcium sulphate, CaH_2O_2 will be calcium hydrate. Water, in fact, might be written HIO , the first H being replaceable by an element according to the usual law of atomicity. A portion of the hydrogen must always remain, or the body would become an oxide. Thus, if potassium were to act upon water, according to the reaction $\text{K}_2 + \text{H}_2\text{O} = \text{K}_2\text{O} + \text{H}_2$, or calcium were to act according to the reaction



the bodies produced are not hydrates, but oxides.

The oxides, which, by addition of water, produce hydrates, are called **ANHYDRIDES**. By subtracting H_2O from any hydrate we may reproduce the corresponding anhydride. Many of the common acids may in this way furnish anhydrides, some of which are interesting bodies.

Sulphuric acid.

Sulphuric anhydride.



In the same manner the student may deduce:

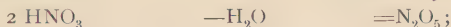
Sulphurous anhydride SO_2 from Sulphurous acid.

Carbonic " CO_2 from Carbonic "

If the acid contains but one atom of hydrogen, we must, of course, double the formula before subtracting. Hence

Nitric acid

Nitric anhydride



similarly we deduce

Nitrous anhydride N_2O_3 from Nitrous acid.

Phosphoric " P_2O_5 " Phosphoric "

We may proceed in the same way with the metallic hydrates,

Calcium hydrate



Calcium anhydride



As before when the hydrate contains but one atom of hydrogen, we double the formula.

Potassium hydrate



Potassium anhydride



These metallic anhydrides are often called BASES.

Hydrogen Dioxide, H_2O_2 . This body, sometimes called oxygenated water, is prepared by liberating oxygen in the presence of water, as when barium dioxide is dissolved in dilute hydrochloric acid ;



A dilute solution may be easily obtained in this way, but the concentrated liquid is very difficult to prepare. It is a colorless, transparent oily liquid, nearly one-half heavier than water ; it is without odor, has a bitter taste, blisters the skin, and bleaches organic colors. It is decomposed by heat and by many chemical substances, often explosively. It dissolves in ether, and the solution has been used in medicine, and for bleaching the hair.

NITROGEN, N.

ATOMIC WEIGHT 14.

DENSITY 14.

Sources. Constitutes about four-fifths of air and occurs in many animal and vegetable tissues, especially in those performing the higher functions. It also occurs in the mineral kingdom in the form of the sodium and potassium nitrates. The name means "producer of nitre." Discovered by Rutherford in 1772.

Preparation. The simplest method of preparing nitrogen is to burn out the oxygen from a limited amount of air. A substance of active combustible qualities like phosphorus or sodium is required for this purpose, as common combustibles, like coal, will be extinguished long before the oxygen is exhausted.

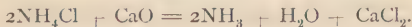
Exp. A piece of phosphorus is placed on a small metal basin or block of wood floating upon water, and being lighted, a bell jar or wide mouth bottle is placed over it. The white fumes of phosphoric anhydride P_2O_5 soon fill the jar, and are absorbed by the water, which rises to about one-third the height of the jar. The nitrogen so obtained is impure, but shows well all the properties of the gas. The pure gas may be made by the action of chlorine upon ammonia, or by heating a mixture of potassium nitrite and ammonium chloride.

Properties. Nitrogen is a gas the properties of which are generally negative. It has no color, taste, or smell, does not, under ordinary conditions, burn nor support combustion. It is not poisonous, but will not support life. From this absence of striking properties no interesting experiments can be performed with it. At high temperature, and under the influence of electric sparks, it will enter in combination with a number of elements. In this way compounds may be formed with oxygen, boron, silicon and carbon, and with a number of metals, such as hydrogen, magnesium and titanium. It is a little lighter than air, 47 cubic inches, weighing 14 grains; 1 litre weighs 1.25 grains. It can be liquified only by intense cold and pressure.

Gen. Chem. Rel. Nitrogen is generally a pentad, sometimes it acts as a triad or even as a monad. Although in the free state it shows so little affinity, it forms by indirect means, compounds with most of the elements. It is an essential ingredient of all the higher tissues of animals, and exists also in vegetable structures, but not so abundantly. In many of its compounds it appears to be held with a low affinity, and these compounds are apt to decompose. Most of the powerful explosives now in use, gun cotton and nitroglycerine, for instance, owe their qualities to the nitrogen present. Many of its compounds are of great practical importance.

Tests. Free nitrogen is recognized by its chemical indifference. In combination it is usually detected by being converted into ammonia, by the action of alkalies.

Ammonia NH_3 . This substance has been known in some forms from a remote period. The name is derived from the temple Jupiter Ammon, in the Libyan desert, which was a point at which the materials for making one of the compounds of ammonia were collected. Ammonia exists in the air in very small quantities. It is given off in the decomposition of organic matter, especially animal remains, and was originally derived from refuse of this kind. It is also produced by the action of hydrogen on nitric acid. The great source at the present time is the water which has been used for washing the common illuminating gas. The distillation of coal gives rise to considerable ammonia, and this is removed by passing the gas through water. The so-called *ammoniacal liquor* is neutralized with an acid, and the resulting compound properly purified. Ordinarily, hydrochloric acid is used, and the reaction is $\text{NH}_3 + \text{HCl} = \text{NH}_4\text{Cl}$. For the preparation of ammonia some compound containing it is decomposed by an alkali. The usual method is by treating a mixture of ammonium chloride NH_4Cl with lime; the reaction is



By passing the gas over dry lime, the water is absorbed and the pure NH_3 is collected.

Ammonia is a colorless gas, of a pungent odor. It is absorbed by water in large amounts; one pint absorbing 700 pints of gas, and increasing fifty per cent. in volume. This solution exhibits most of the properties of the gas, and is much used under the name of *aqua ammoniæ*, or solution of ammonia.

Ammonia burns, but not easily. It is lighter than air. 1 litre weighs 0.76 gm. 47 cubic inches weigh 8.5 grains. It contains one volume of N and three volumes of H, condensed to two volumes. At a temperature of -40°F . (-40°C .), or under a pressure of 100 lbs. to the square inch, it condenses to a colorless liquid. This liquid, of course, evaporates rapidly when the pressure is removed, and produces great cold which fact has been made use of in machines for making ice.

Gen. Chem. Rel. The important property of ammonia is the strongly alkaline and basic power of its solution in water. This solution has chemical characters so much like those of potassa and soda, that it is now generally believed that the ammonia and water have combined to form a compound, which is analogous in composition to the alkalis proper. Common caustic potassa and caustic soda are known as hydrates, and have the composition KHO and NaHO . If we add NH_3 to H_2O we get a body which can be written NH_4HO , in which the NH_4 takes the part of the K or Na. To indicate that this molecule, NH_4 , is like a metal in its character, it is called *ammonium*, and its compounds are called ammonium salts. The action of HCl upon NH_3 is supposed to produce NH_4Cl , and the result of the action is now called ammonium chloride, not muriate of ammonia, which was the old name. The series of ammonium compounds is quite similar to those of potassium and sodium, and will be described in connection with those elements and the ammonium theory still further explained.

Exp. The affinity of ammonia for water may be easily shown by filling a flask with the gas and inverting it in water. The ammonia will be rapidly absorbed, and the water will rise into the vacuum thus produced.

Tests. Ammonia is recognized by its odor, its alkaline reaction and the white cloud of NH_4Cl produced when it comes in contact with vapors of hydrochloric acid. The most delicate test for its presence, either free or in combination, is Nessler's reagent, a solution made by mixing HgCl_2 , KI , and KHO or NaHO . The liquid so produced, produces with very minute quantities of ammonia a yellow color or a yellowish red precipitate. One part of ammonia in fifty million parts of water can be easily recognized.

Nitrogen oxides. Five compounds of nitrogen and oxygen have been obtained in the free state. These are:

N_2O	Nitrous oxide. Laughing gas.
NO	Nitric oxide, (often written N_2O_2).
N_2O_3	Nitrous anhydride.
NO_2	Nitrogen peroxide, (often written N_2O_4).
N_2O_5	Nitric anhydride.

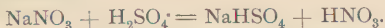
The system of names given to these compounds is in some confusion, owing partly to the fact that they were formerly assigned a somewhat different composition, and partly to the practice of some writers in using for some of the compounds a doubled formula. Thus, NO is often written N_2O_2 and called nitrogen dioxide, because of the O_2 present. NO_2 is written N_2O_4 and called nitrogen tetroxide. The names and formulæ given above are in the main preferable.

Nitrogen and oxygen combine to a limited extent under the influence of very high temperature, especially when produced by an electrical discharge, generally producing one of the higher oxides. The common source of the different oxides is, however, by decomposition of nitrates and of nitric acid, which latter body it will be convenient to describe first.

Nitric Acid. *Aqua fortis*, HNO_3 . This body has been known in the free state for a long time, and two of its salts, KNO_3 and $NaNO_3$, are found in large amounts as minerals; the former in India, the latter in Peru and Chili. The usual method of obtaining the acid is by the action of strong sulphuric acid upon the nitrates. For commercial purposes sodium nitrate is used, being the cheaper salt. The reaction is



The complete action requires a high temperature. For a laboratory experiment it is preferable to use less sodium nitrate, and the reaction will be



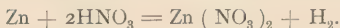
Exp. Mix in a retort equal weights of powdered sodium or potassium nitrate, connect the retort with a receiver set in cold water, and apply to the mixture a gentle heat. The nitric acid distills over, and will condense in the receiver as a slightly yellow fuming liquid. The experiment is best performed in a well ventilated place. No corks or gum tubing should be used in connecting the apparatus.

Nitric acid thus obtained has the composition HNO_3 , but it is difficult to prevent it absorbing water, and it is generally yellow from slight decomposition: when quite pure it is colorless. The

ordinary acid has the composition $2\text{H}_2\text{O} + \text{HNO}_3$. It is a strongly acid liquid, decomposing in the light, highly corrosive and poisonous, and of very active chemical qualities. Its special value in chemistry is its high oxidizing qualities. About one-half the oxygen which it contains is available, and is given up to a great variety of substances. Some variation occurs in the action of the acid, but the effect is in most cases represented as follows:



The O_3 is the available oxygen; the NO will be hereafter described as nitric oxide, and its escape is one of the most characteristic evidences of the action of the acid. With some metals of low affinity the acid acts simply by losing hydrogen and taking up the other metal. Thus:

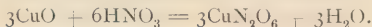


The evolved hydrogen, however, attacks another portion of nitric acid, and forms ammonia. This effect is made use of in the detection and estimation of nitrates, especially in drinking water. The action can be easily observed by making a mixture of zinc and dilute sulphuric acid, and adding nitric acid. The hydrogen which is copiously evolved when the sulphuric acid alone is present, slowly diminishes upon adding the nitric acid, and may even cease to come off. After a time a perceptible amount of ammonia will be formed. The action is even better produced when we dissolve aluminium foil in a solution of caustic alkali, mixed with a small quantity of a nitrate.

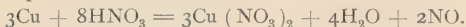
With metals like copper, silver and mercury, which do not easily displace hydrogen, the action of nitric acid is highly characteristic. It exerts the special oxidizing action above referred to. Two molecules of the acid give up three of oxygen, and form with the metal the compounds that are called basic oxides or bases. We may suppose, in the case of copper, the following to take place:



The NO escapes as a gas; the 3CuO immediately act upon and neutralize six additional molecules of nitric acid,



The complete reaction, therefore, is



Similar results are obtained with many other metals.

A third action of nitric acid is its power of forming substitution compounds. The general nature of these bodies belongs to organic chemistry, but a single illustrative reaction may be here given. When benzole C_6H_6 is treated with strong nitric acid,

one atom of hydrogen is removed, and one molecule of NO_2 put in its place. We have



and the body so formed is called *Nitrobenzole*.

A mixture of nitric and sulphuric acids is often used for such effects.

With many organic bodies nitric acid simply acts by adding oxygen, and often gives them a bright yellow color.

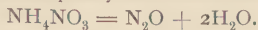
Very strong nitric acid fails to act upon some substances, which are readily attacked by the more dilute forms.

Tests. The strong acid may be recognized by its odor, and its power of producing yellow stains on organic matter.

Free nitric acid produces the following marked reactions:

Morphia and brucia are colored red; copper is dissolved with the production of red fumes of NO_2 . When the acid is in combination, a few drops of sulphuric acid should be added. This will liberate the nitric acid and develop the tests.

Nitrous oxide, N_2O , laughing gas, often called nitrogen monoxide. This body was discovered by Priestley in 1776. It is best obtained from ammonium nitrate, which, when carefully heated, decomposes completely into nitrous oxide and steam.



Nitrous oxide is a colorless, odorless gas, with a somewhat sweetish taste. It can be breathed for a short time without injury, and produces a transient intoxication. When the gas is inhaled in quantity, a short insensibility is produced, during which brief operations such as teeth extractions and opening of abscesses may be performed. It supports the burning of ordinary combustibles almost as well as oxygen. It is somewhat soluble in water. At a pressure of fifty atmospheres it becomes a colorless liquid, and is now sold in this form, compressed in strong metal cylinders.

Nitric oxide, NO , often called nitrogen dioxide, and written N_2O_2 . This gas is the usual product of the action of nitric acid as an oxidizer. It is conveniently prepared by the action of nitric acid on copper. The reaction is



NO is a colorless gas, very difficult to reduce to the liquid form. Its most marked property is that as soon as it is brought in contact with oxygen, it instantly absorbs one atom, becoming NO_2 , and turning brownish red. This body NO_2 parts with the second atom of oxygen rather easily, and leaves the NO ready to absorb more oxygen. In consequence of this property, nitric oxide is

used as a sort of carrier of oxygen, particularly in the manufacture of sulphuric acid. In the ordinary method of making nitric oxide, a portion is always converted into NO_2 ; this can be removed by collecting the gas over water, which absorbs the NO_2 .

Nitrogen peroxide, or nitric peroxide NO_2 . This body has various names, owing to uncertainty in its chemical relations. It has been called nitrogen tetroxide, (being written N_2O_4 by some chemists), and hyponitric acid, and, also by other less common names. The proper name would be nitrogen dioxide, to correspond to the formula NO_2 , but as this name has, unfortunately, and very unwisely, been given by some writers to nitric oxide it is probable that the rather unsatisfactory name, nitric peroxide, will have to be retained.

NO_2 is a brownish red gas of which the color depends somewhat on the temperature; at about 14°F , (-10°C), it condenses to a liquid. It is readily absorbed by water, which decomposes it, producing various compounds, according to the proportions used.

The following experiments illustrate the general relation of the nitrogen oxides:

1. Mix in a glass bottle or retort, some copper turnings with moderately strong nitric acid. The evolved gas will have a red color, and consist of a variable mixture of N_2O , NO and NO_2 . Pass the gas into a jar of water inverted over the pneumatic trough, and it will be found that a colorless gas is collected. This is because the NO_2 is absorbed by the water.

2. Allow a few bubbles of air to pass up into the jar, a red color is at once produced, which soon disappears, the water rising slightly in the jar. This effect is due to the conversion of the NO into NO_2 , and the subsequent absorption of the latter by the water. The experiment may be repeated until all the NO is oxidized, and the residual gas is a mixture of N_2O and N .

3. If pure nitric acid and pure oxygen be used in this experiment, the relation by volume will be easily shown. Two pints of nitric oxide contain one pint of N and one of O . If mixed with an additional pint of oxygen, NO_2 will be formed, and complete condensation will ensue.

Nitrous anhydride N_2O_3 and nitric anhydride N_2O_5 are unimportant. Nitrous acid HNO_2 is also of little importance, but a few of the nitrites are of considerable chemical interest. Ammonium nitrite occurs in rain water, and other nitrites are found in river and spring water. They act both as oxidizing and reducing agents.

Air. The atmosphere is an intimate *mixture* of about four volumes of nitrogen with one volume of oxygen. It surrounds

the world like a shell or casing, and extends upwards to a height which has been variously estimated at from 45 to 200 miles. It is known not to be a compound by several tests, among which are:

1st. It turns brown an alkaline solution of pyrogallin, and reddens nitric oxide, both of which characters belong only to free oxygen.

2nd. Air which is dissolved by water has its nitrogen and oxygen in a proportion somewhat different from that in which they occur in undissolved air. A chemical compound would have the same composition in both cases.

3rd. The composition is not perfectly constant and is not in exact proportions either by weight or volume, though its composition approaches very closely that required by the formula N_4O .

The fact that this mixture of gases varies so little in different places and under different circumstances is due to an action called diffusion, which is seen in all gases, in nearly all liquids, and would probably be noticed in some solids, if high pressures should be used. All gases mingle with each other so that sooner or later they produce a uniform mixture, in spite of the influence of gravity. The rate of mixture is dependent on the density of the gas. It is expressed mathematically by saying that the rate of diffusion is *inversely* proportional to the *square-root* of the densities. Suppose two gases have densities of 1 and 16 respectively; the square-root of the numbers will be 1 and 4, and then by inversion we find that the lighter gas will diffuse with 4 times the rapidity of the heavier. Experiments on diffusion are usually shown with hydrogen, because, being very light, it shows the effect quickly. Such experiments belong to physics rather than chemistry. The rapidity with which the odor of coal-gas penetrates through a room, when a leak in a pipe occurs, is an instance of diffusion. If the gas were to obey the usual laws of gravity, it would rise to the ceiling, but it really penetrates to all parts almost the same as if it were escaping into an empty space. In the same way, although nitrogen has a density of 14 and oxygen of 16, they do not form separate layers, but are uniformly and permanently mingled.

Ordinary air contains small quantities of many other bodies besides nitrogen and oxygen. It always contains water, carbonic acid and ammonia, and frequently some compounds of nitrogen and oxygen, and also ozone. Besides these we have dust and the products of animal and vegetable decomposition. The study of impurities of air has received much attention of late years, more especially in view of the theory that many diseases are due to living organisms or germs which exist in the air. For these investigations the microscope has been necessary. Some of

the more important of these results are given under the organic chemistry.

The approximate composition of air was first demonstrated by Lavoisier in 1777.

The accurate analysis varies somewhat: the following may be taken as a fair average:

Oxygen,	20.61.
Nitrogen,	77.95.
Carbon dioxide,	.04.
Water	1.40.

Traces of ammonia, nitric acid and marsh gas, (CH_4), and, in towns, sulphur compounds.

The uniformity of the composition of air is assisted by the winds and currents which are continually agitating it. Its chemical properties are those of oxygen, but in a much diminished degree, on account of the extensive dilution with nitrogen. It was formerly taken as a standard for the specific gravity of gases, but hydrogen is now preferred. 100 cubic inches weigh 30.93 grains. 1 litre weighs 1.29 grammes. 13 cubic feet weigh about 1 lb. At the level of the sea the pressure is, ordinarily, about 15 lbs. and will sustain a column of mercury 760 millimetres, or 30 inches in height. Water in its natural condition always contains some air in solution.

The capacity of air for holding moisture increases rapidly as the temperature rises. The dryness or dampness of the atmosphere is not due to the actual quantity of moisture in it, but to the amount in proportion to what the air can take up. A cubic foot of air at 30°F can absorb about 2 grains of water; if it contains a grain and three-quarters it will, therefore, be nearly saturated and seem damp; if the temperature rise to 80°F the capacity for moisture will rise to 11 grains, and under these conditions it would seem dry if it held three grains of moisture, because, although the amount for moisture is nearly doubled, its capacity for moisture has increased over five times. The nearness of the air to saturation is called the **RELATIVE HUMIDITY**. Air saturated with water has a relative humidity of 100; if half saturated the relative humidity is 50, and so on. When the temperature falls the moisture separates to a greater or less extent, and we have fog, rain or dew, and if the temperature get below the freezing point we have snow or frost.

Changes produced in the atmosphere. The changes which affect the chemical composition of the air are important. The respiration of animals and the processes of combustion are continually removing oxygen and introducing water, carbonic

acid and more or less organic matter. The decay of animal and vegetable substances introduces various gases, especially ammonia and hydrogen sulphide; and the burning of coal and coal gas furnishes sulphurous and sulphuric acid. The dust which is always floating in the air contains a great variety of substances, living and dead, and varies with the locality. The continued removal of oxygen is counterbalanced by the respiration of plants which, under the influence of light, decompose the carbonic acid, retaining the carbon and giving off the oxygen, especially at the under surface of the leaves. In this way the two great divisions of organic nature sustain each other. This fact is well shown in the construction of the ordinary aquarium, in which animal and plant life are maintained for a long time without renewal of the water. The nitrogen of atmosphere is very little affected. The ammonia and other gases are gradually oxidized or absorbed by the soil and plants, and washed out by the rains. The organic matter also oxidizes, and ozone is supposed to be especially active in this respect.

Ventilation. When animals are compelled to breathe air in closed space it becomes, sooner or later, by the removal of oxygen, incapable of supporting their life, and by the introduction of organic matter it becomes an active agent of disease. As all buildings interfere with the free circulation of the air, the problem of ventilation or proper renewal of the air is a very important one. The products of respiration and combustion are usually lighter, because warmer than the air around, and tend to rise, and the simplest systems of ventilation take advantage of this fact, by arrangements which allow the foul air to escape at the top of the room, and fresh air to enter at the bottom. Unfortunately, walls and windows are usually cold; they chill this foul air and interfere seriously with its upward movement. Practically, good ventilation can not be accomplished without some active mechanical assistance, such as a fan-blower or the draft of a chimney.

CARBON, C.

ATOMIC WEIGHT, 12. DENSITY, IN THE STATE OF GAS,
UNDETERMINED.

Sources. Occurs very abundantly in nature, both in the free state and in combination. It is so constant a component of organic bodies that organic chemistry has been rather romantically called the chemistry of the compounds of carbon. In the tis-

without
 sues of animals and plants it exists in union with hydrogen, oxygen and nitrogen. The various forms of coal and graphite, and certain carbonates, especially of calcium and magnesium, are abundant minerals. Carbon is remarkable for presenting itself under modifications so different that we would hardly suppose them to be of the same composition. Such modifications of properties ~~with~~ change of composition are called *allotropic*. The allotropic forms of carbon are:

Amorphous Carbon, of which lampblack and charcoal are examples.

Graphite or plumbago, which is not perfectly pure, and is probably a very old form of coal.

Diamond, which is often chemically pure, but is also found in inferior conditions.

Some properties are common to all these forms. They are all insoluble in all liquids, and infusible and unacted upon by acids and alkalies, or by the air at ordinary temperatures. Heated strongly in air or oxygen, they burn, producing CO or CO₂.

A number of impure forms of carbon are also known. The special properties and origin of the forms, pure and impure, need only brief description.

Lampblack is the deposit from smoky flames. It is a soft black substance, used for printing inks and colors. It generally contains hydrogen.

Wood Charcoal is obtained by heating wood out of contact of air. It contains hydrogen and the mineral substances of the wood.

Animal Charcoal is obtained by charring blood, bones and other animal tissues. It is coarse black powder.

Wood and animal charcoals have great powers of absorption; the former for gases, the latter for organic matters, especially colors and bitter principles.

• Exp. If a piece of wood charcoal be weighted so as to sink in some water in a tall jar, and the jar then be placed under the receiver of an air pump, and the air exhausted, large quantities of gas, principally nitrogen, oxygen and carbon dioxide, will escape from the charcoal. This property of wood charcoal explains its use as a deoderizer. Gases containing hydrogen, sulphur or phosphorus, are generally entirely burned up or decomposed when absorbed in this way.

Exp. If a solution of some organic color, such as litmus or cochineal, be filtered through animal charcoal, the color will be partly or wholly removed. Bitter principles such as strychnia, or the bitter of hops, will also be removed. Animal charcoal is

extensively used for decolorization of syrups and vegetable infusions generally.

Graphite, called also plumbago, and black lead, is destitute of any absorbent properties, and is used for lead pencils, and for crucibles.

Diamond is a crystalline form of carbon, and has been produced artificially, though in very minute form. Its origin is unknown, and it is generally found in soil which has been transported by water. It is the hardest substance known, and in the impure and discolored forms it has been used with great advantage for the drilling and cutting of stone.

The secondary properties, such as specific gravity, color and hardness are different in the various forms of carbon.

The impure forms of carbon are the varieties of coal.

Coal has been formed from vegetable matter by a slow process of decay, mostly under water, by which the hydrogen and oxygen are, in great part, removed, and the carbon by pressure made compact. Bituminous or soft coals are by such action produced. They always contain hydrogen and oxygen, and when heated produce a variety of gases which have illuminating power, and constitute ordinary coal gas. Remains and impressions of plants are found in such coal. Coke is the residue after heating the coal.

Anthracite coal is much harder and has very little hydrogen. It yields no gas on heating. It is practically a compact coke. The anthracite coal fields of Pennsylvania are the most valuable in the world; very few deposits like them being known. In the north-western part of the United States, coal of comparatively recent origin occurs.

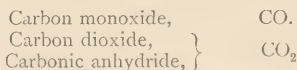
Gen. Chem. Rel. Carbon is a tetrad, and combines many elements, especially with the non-metals. Its compounds with the metals have not been extensively studied, except those with hydrogen, which are quite numerous, and with iron. Cast iron is substantially a carbide.

Tests. Carbon in the free condition is easily recognized by its infusibility and combustibility, and by producing carbonic acid. Most of its complicated compounds yield a black residue of charcoal, when strongly heated, or mixed with strong sulphuric acid.

Compounds of Carbon with Hydrogen. Hydrogen and carbon combine in many proportions, forming a great variety of bodies. Some of the substances are produced by elaborate chemical operations, some by the natural processes of decay, many important ones by the action of heat on organic substances, especially coal and wood.

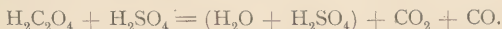
Coal gas. When bituminous coal is heated in a closed vessel large amounts of gas, (about 5 cubic feet to the pound, in good samples), are given off. This gas is contaminated with tar, ammonia and sulphur compounds. It deposits its tar by cooling; the ammonia is removed by water, and the sulphur by lime, and thus purified, it constitutes illuminating gas, which is a variable mixture of hydrogen, marsh gas CH_4 , olefiant gas C_2H_4 , and other gases. Further accounts of these will be found under the chemistry of organic bodies.

Compounds of Carbon with Oxygen. The important ones are:



Carbon Monoxide. Carbonic oxide CO . This body is produced when carbon is burned in a deficient supply of air, as in stoves with defective draft, and in the large furnaces for reducing and working iron, when an excess of fuel is purposely maintained. When steam is thrown upon burning coal, the reaction, $\text{H}_2\text{O} + \text{C} = \text{CO} + \text{H}_2$ occurs, and the resulting mixture is available as a gaseous fuel, or may be impregnated with vapors of benzine or gasoline, and used as a source of light. In making carbon monoxide for experimental process these methods are unsuitable, and application is made of the fact that the action of sulphuric acid upon oxalic acid, or upon potassium ferrocyanide, gives rise to the gas.

Exp. Mix in a rather large flask, some crystallized oxalic acid with its own bulk of sulphuric acid, and heat carefully. The mass will soon begin to foam and give off a mixture of CO and CO_2 . By passing this through lime water or caustic soda, pure carbon monoxide may be obtained. The reaction is



The sulphuric acid and water form a sort of combination.

If potassium ferrocyanide be used in the proportion of about 1 part by weight to 13 parts of sulphuric acid, and 1 of water, a large volume of the pure gas may be obtained. The reaction is a complicated one, and the mixture swells considerably, so that a large flask must be used.

Properties. Carbon monoxide is a colorless, odorless, tasteless gas, of decidedly narcotic-poisonous properties. It is a little lighter than air. It burns easily with a clear blue flame. It is an unsaturated molecule, and will combine with chlorine and some other elements.

Carbon Dioxide. Carbonic Anhydride, CO_2 , generally

occurs in union with H_2O , forming H_2CO_3 carbonic acid, a very abundant substance, occurring in air and water. Some of its salts, especially calcium and magnesium carbonates, are common minerals.

Carbonic Acid is produced in a great variety of ways.

1. By the respiration of animals.
2. By ordinary combustions.
3. By fermentation and decay.
4. By decomposition of carbonate, either by heat or acids.

The last method is made available for its production on the small scale.

Exp. Put some fragments of marble or chalk, (CaCO_3), into a retort or gas bottle, and add some moderately strong hydrochloric acid. The escape of gas occurs at once and it may be collected by downward displacement or over water. Sulphuric acid does not answer so well, as it soon forms an insoluble mass in the bottle.

The reaction is



It is unimportant whether we regard the water and CO_2 as separate or united. By passing the escaping gas over dry calcium chloride, or strong sulphuric acid, the pure CO_2 may be collected. Its properties are substantially those of H_2CO_3 .

Properties. Carbonic acid is a colorless gas of a somewhat sharp taste. It is soluble at ordinary pressure in its own bulk of water, and the solubility is increased in regular proportion to the pressure. It is about fifty per cent. heavier than air, and may, therefore, be easily collected by running the delivery tube to the bottom of the jar. 1 litre weighs 2.07 grms. 47 cubic inches weigh 22 grs. It can be liquified by a pressure of 800 lbs. to the inch, and freezes at -70°F . (-56°C). It does not support animal life, nor ordinary combustion, but bodies of high affinity, if already in active combustion, will decompose it, and continue to burn. In this way red hot coal will produce the following reaction: $\text{C} + \text{CO}_2 = 2\text{CO}$; which accounts for the production of carbon monoxide in ordinary stoves.

Exp. A lighted taper put into the gas is instantly extinguished, but a slip of ignited magnesium will continue to burn and deposit carbon. The reaction is $\text{Mg}_2 + \text{CO}_2 = 2\text{MgO} + \text{C}$.

Exp. Collect the gas in a tube over water, introduce a small piece of caustic soda, and quickly cork the tube. After a few moments shaking, open the tube with the mouth under water, when the rise of the water will indicate the absorption of the gas.

Exp. Lime water is instantly rendered turbid by the gas from the formation of insoluble calcium carbonate, but if the quantity of the carbonic acid is in excess, the precipitate will be redissolved, producing what is ordinarily known as a hard or limestone water. This action is more fully described in connection with the calcium salts. If the clear solution produced in this experiment be boiled, the excess of CO_2 will be expelled, and the precipitate will be reproduced.

Exp. Soap bubbles blown with the gas sink rapidly in the air, and if blown with ordinary air will float in a jar of the gas.

Exp. A large light vessel being counterpoised on a delicate balance will be thrown decidedly out of balance by substituting carbonic acid for the contained air. By inverting the vessel the gas will escape, and the balance be restored.

Exp. Carbonic acid may be easily poured from one vessel to another, and if a lighted taper be put at the bottom of a tall beaker, it will be extinguished by the gas falling over it.

In accordance with the law given on page 47, carbonic acid has low diffusive power, and hence a tendency to accumulate at low levels, if produced in large amounts. It is found in undue proportions at the bottom of mine shafts, and in fermenting vats, and cases of suffocation often occur in these places. The usual method of determining whether such places are safe to enter, is by lowering a lighted candle, if this continues to burn vigorously, the air is probably safe; if it burns feebly or is extinguished, the air is too rich in carbonic acid.

Gen. Chem. Rel. Carbonic acid may be considered a feeble chemical agent when compared to such bodies as nitric and sulphuric acids, but its continual presence in air and in water makes it one of the important agents in the slow changes which occur in nature. Assisted by the action of frost, it breaks down, and renders soluble many kinds of rocks, and converts them into soils. Its power of increasing the solvent action of water makes that liquid a very active agent in geological changes. The chemical activity of its solution in water is increased by pressure, and water which is charged with it at great depths, often emerges at the surface holding much mineral matter. The release of pressure causes the gas to escape and gives an effervescing water. The mineral matter is at the same time deposited. The ordinary effervescing soda water is an artificial solution of the gas under pressure. Fermenting liquids owe their effervescence to the same cause, the retention of the gas under pressure, and its escape when the pressure is released. Its relation to plant life is very important. Under the influence of light, plants decompose it, the carbon being absorbed, and the oxygen given off.

Carbonic acid forms a series of salts, called the carbonates, most of which are insoluble in pure water. The monad metals form two salts. Potassium gives us, for instance,

KHCO_3 Acid potassium carbonate.

K_2CO_3 Potassium carbonate.

Dyad metals give but one salt, as

CaCO_3 Calcium carbonate.

The carbonates are decomposed by almost all acids, and generally by heat.

Tests. Carbonic acid is easily recognized by rendering turbid a solution of calcium hydrate, (lime water), or barium hydrate, (baryta water). It turns litmus to a wine-red, the blue color being restored on boiling; solution of cochineal is not effected.

Combustion and the structure of flame. The elements carbon, hydrogen, oxygen and nitrogen, and their compounds form by far the greater part of all the material objects around us, and are the especial elements of the tissues of animals and plants, from which our fuel and illuminating agents are directly or indirectly derived. The process of burning is much the same as the process of decay. It is the absorption of oxygen and the formation of carbonic acid, water and free nitrogen, or sometimes of ammonia and nitric acid. The phenomenon of flame attracted the attention of the most ancient investigators, and it was by them considered a form of matter. We know now that ordinary flame is a process; it is gas of some kind, in the act of uniting with the oxygen of the air, the operation being attended with the production of light, heat and other forms of force.

Formerly the terms "combustible" and "supporter of combustion" were much used; carbon, phosphorus and hydrogen being called combustible elements, oxygen and chlorine supporters of combustion. This distinction is now abandoned; the action is a mutual one, and the supporter of combustion may easily be made the combustible. True flames may be produced in which none of the ordinary agents are used.

If we examine common gas or candle flame we find it consists of three parts:

1. An inner space of a blue color.
2. A shell of brightly luminous particles.
3. A fringe of feebly luminous particles.

The flame of a candle or other body burning free in the air is generally pointed or conical, due to drafts of air which strike the side of the flame and rise, drawing in toward the centre. When, as in our gas burners, the burning body is supplied under

pressure, the form of the flame is different, but the different flames are still conveniently called *cones*.

The inner cone is the point at which the gas that is burning is produced or escapes. In the flame of a candle or of coal oil the gas is the result of a destructive distillation of the fat or oil. This gas contains carbon and hydrogen. At its outer edge it meets the air; most of the hydrogen is converted into water, the carbon is set free in solid but finely divided condition, in union with some hydrogen, and this solid is intensely heated by the combustion of the hydrogen. This is the source of the light, and forms the second cone. The finely divided matter passes outward and gradually burns, producing the feeble fringe of light, which is the third cone. It is obvious that with bodies which are deficient in carbon, or which are burned in a supply of oxygen sufficient to consume the carbon before it can be set free, very little light will be produced; on the other hand, if the quantity of carbon is large, the flame will not be able to heat it above a red heat, and the supply of oxygen may not be sufficient to burn it up, and we then have a lurid, smoky flame.

The following experiments show the general nature of flame, and the manner of studying it:

1. The flame of pure hydrogen gives out very little light. If solid particles of charcoal dust, platinum wire, etc., be introduced, they will become luminous.
2. If a small glass tube be introduced into the centre cone, a quantity of the unburnt gas will pass out, and may be burned at the other end of the tube. The temperature of this interior cone is very low.
3. Alcohol, which contains very little carbon, burns without much light; turpentine, which contains much carbon, burns with a red flame and smoke. By making a mixture of the two a pretty good flame may be obtained.
4. A piece of paper or card being held for a moment low down in a spirit lamp flame, will be charred in a ring, showing that the interior of the flame is not burning.
5. Anything which cools the carbon down below its burning point will cause it to deposit in the solid form; hence the formation of soot or lampblack, when flames come in contact with cold surfaces.
6. If a flame be suddenly cooled, as by the introduction of a coil of wire, or a sheet of wire gauze, the combustion will cease and the mixture of gas and air will escape. This can be easily shown by putting a piece of wire gauze across a gas flame, when it will be found that the flame will stop at the gauze, but a

combustible mixture of gas and air will pass through it. Similarly the gas may be lighted above the gauze, and the flame will not run back. If the gauze becomes hot, the flame will pass through.

This principle is made use of in the Davy's safety lamp, for preventing explosions in mines. It consists of a lamp arranged so that no air or gas can get in except through fine gauze. If an explosive mixture finds its way to the flame its combustion is limited to the interior of the lamp, at least for a time.

If common coal gas be mixed with air, it will burn with a non-luminous, smokeless flame, and such lamps are now used very largely. In the simplest form, the Bunsen burner, the air is drawn in through openings at the bottom. A great variety of these lamps are now made, and Mr. Fletcher of England, has brought the use of gaseous fuel to high perfection, so that it bids fair to come into almost exclusive use.

When a current of air is driven into a flame its temperature is increased. This is the cause of the efficacy of the mouth blow-pipe, and of blast-lamps.

When mixtures of gas and air are ignited, combustion may occur through the entire mass at once. This constitutes an explosion.

Cyanogen, CN. Cyanogen is electro-negative, and in its chemical relations resembles such elements as Cl, Br and I. It forms compounds called *cyanides*. In all of these it acts as a monad; thus we have hydrogen cyanide HCN, potassium cyanide KCN. Dyad metals require, of course, two molecules of cyanogen. Calcium cyanide is CaC_2N_2 or $\text{Ca}(\text{CN})_2$. The symbol Cy is often used in formulæ instead of the symbol CN. We write HCy instead of HCN, KCy instead of KCN, CaCy_2 instead of the formula for calcium cyanide just given. The chemistry of cyanogen is more conveniently considered in connection with organic compounds.

SULPHUR, S.

ATOMIC WEIGHT 32. DENSITY 32.

Sources. Occurs native, i. e. in the free state, in volcanic regions, also in combination with the metals, forming sulphides and sulphates, and in animal and vegetable structures.

Preparation Commercial sulphur is prepared by melting or distilling the native sulphur, or some of the sulphides. It pre-

sents itself in two forms: *roll sulphur* or *brimstone*, made by casting the melted sulphur in moulds, and *flowers of sulphur*, made by condensing the distilled sulphur in a cool chamber. *Lac sulphuris*, or milk of sulphur, is a finely divided medicinal form, obtained by dissolving common sulphur in milk of lime, and precipitating by an acid.

Properties. Sulphur assumes several allotropic forms, varying especially in color and solubility; it is dimorphous, i. e., crystallizes in two forms—octahedral and prismatic. Ordinarily, sulphur is brittle, yellow and soluble in carbon disulphide; but by being suddenly cooled from near its boiling point, it becomes plastic, dark colored, and insoluble. All varieties are insoluble in water, highly combustible; fusible at about 250°F. (121°C.), and boiling at 836°F. (447°C.). It is a non-conductor of electricity, and becomes highly electrical by friction.

Exp. A small quantity of sulphur is placed in a flask, and heated slowly. It melts to a thin, amber colored liquid. On continuing the heat, the liquid gradually becomes thick, and at about 450°F. (232°C.), it is so tenacious that it can scarcely be poured out of the vessel. Heated still further it becomes thin, and finally boils. Just before the sulphur begins to boil, pour it into cold water, when it will form dark brown semi-elastic masses. If the sulphur remaining in the flask be heated to the boiling point, a dark red vapor is produced, in which certain substances, *e. g.* Dutch leaf, burn easily.

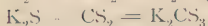
Exp. Dissolve some sulphur in a small quantity of carbon disulphide, and allow the solution to evaporate in the open air. Transparent yellow octahedra will be deposited.

Exp. A few ounces of sulphur are fused in a Hessian crucible, and then allowed to cool until a film of solid forms on the surface. Break a small hole through this and pour out the liquid contents. On breaking the crucible a mass of prismatic crystals will be found.

Gen. Chem. Rel. Sulphur combines with most metals, and often with considerable energy, forming an important group of bodies—the sulphides. In these compounds it is electro-negative and usually diatomic. It also combines with the non-metallic elements, its atomicity varying, according to the usual law, from dyad to hexad. In association with oxygen and chlorine, and, perhaps, with some other elements, it is regarded as electro-positive. In general its compounds are analogous in composition to those of oxygen, and as the oxides of the metals act as bases towards the ordinary acids, so the sulphides of the metals act as bases towards what are called the sulphur acids. Thus we have



Potassium carbonate.



" sulphocarbonate.

In such compounds the sulphur is substituted for the oxygen, atom for atom, and the name is formed by affixing the syllable "sulph" to the name of the acid. A few of those sulphur salts are very important.

Sulphur has many important uses. It is employed in medicine as an alterative, and externally for skin diseases. It is used in the arts for vulcanizing caoutchouc, and in the manufacture of gunpowder. Match sticks are tipped with it to make the friction composition ignite the wood more surely.

The compounds of Sulphur and Hydrogen are two, corresponding to the oxides of hydrogen:

H_2S Hydrogen sulphide, or sulphureted hydrogen,

H_2S_2 Hydrogen disulphide.

Hydrogen sulphide, H_2S . Discovered by Scheele in 1777. This substance is a gas. It exists in solution in some spring waters, also in the emanations from volcanoes, and decomposing animal and vegetable matters. It is sometimes produced by the action of organic matter upon sulphates. Calcium sulphate, CaSO_4 , by losing its oxygen becomes CaS , and this by the action of carbonic acid, yields the gas. $\text{CaS} + \text{H}_2\text{CO}_3 = \text{CaCO}_3 + \text{H}_2\text{S}$. As hydrogen sulphide is much used as a test, it is frequently made in the laboratory, generally by acting upon sulphides with strong acids. Ferrous sulphide and sulphuric acid are much used. $\text{FeS} + \text{H}_2\text{SO}_4 = \text{FeSO}_4 + \text{H}_2\text{S}$.

Exp. Introduce into a gas bottle some ferrous sulphide, in small fragments, and pour over it some dilute sulphuric or hydrochloric acid. The hydrogen sulphide is liberated freely. As it is only slightly heavier than air, and rather soluble in water, it is difficult to collect either by displacement or over water. The experiments with it are usually performed by passing the gas into certain metallic solutions, as described below.

Properties. Hydrogen sulphide is a colorless gas, condensible by moderate cold and pressure. It has a strong odor, like rotten eggs, is easily combustible, burning with a pale blue flame, and producing sulphurous acid, H_2SO_3 . Water at ordinary temperature dissolves about three volumes, acquiring the odor and chemical properties of the gas. The important property of hydrogen sulphide is its power of precipitating many metals, as sulphides. These precipitates being generally distinct in color and highly insoluble, their production is not only a test for the presence of such metals, but also a means of separating them from solution.

Exp. As illustrations, solutions of copper sulphate, mercuric chloride, tarter emetic, and zinc sulphate may be precipitated either by a current of the gas, or by its aqueous solution. The zinc sulphate requires the addition of a few drops of ammonia.

When pure hydrogen sulphide is required it is generally obtained by heating a mixture of hydrochloric acid and antimony sulphide. $\text{Sb}_2\text{S}_3 + 6\text{HCl} = 2\text{SbCl}_3 + 3\text{H}_2\text{S}$.

Hydrogen sulphide is a powerful reducing agent; solution of potassium permanganate is rapidly decolorized by it. A mixture of potassium bichromate and hydrochloric acid is changed into green chromic chloride.

Tests. Hydrogen sulphide is recognized by its smell, and by blackening paper soaked in lead acetate solution. As the gas is a frequent product of decomposition, the leakage of sewer-gas may sometimes be detected by this test.

Hydrogen disulphide, H_2S_2 . This substance may be prepared by boiling together lime and sulphur, forming calcium disulphide, and pouring the solution into dilute hydrochloric acid. The hydrogen disulphide separates as a yellow oily liquid, of a disagreeable odor. It decomposes easily.

Compounds of sulphur with oxygen. Sulphur forms with oxygen a number of acid-forming oxides, or anhydrides, most of which are known only in the hydrated condition, that is, as acids.

Anhydride.		Acid.	Name.
SO		H_2SO_2	Hyposulphurous.
SO ₂		H_2SO_3	Sulphurous.
SO ₃		H_2SO_4	Sulphuric.
S ₂ O ₂	not yet	$\text{H}_2\text{S}_2\text{O}_3$	Thiosulphuric.
S ₂ O ₅		$\text{H}_2\text{S}_2\text{O}_6$	Dithionic.
S ₃ O ₅		$\text{H}_2\text{S}_3\text{O}_6$	Trithionic.
S ₄ O ₅	obtained.	$\text{H}_2\text{S}_4\text{O}_6$	Tetrathionic.
S ₅ O ₅		$\text{H}_2\text{S}_5\text{O}_6$	Pentathionic.

The first member of this series is the true hyposulphurous acid, the commercial hyposulphites being really salts of thiosulphuric acid, hence properly called thiosulphates. Most of the members of the above series are of limited interest. Only a few need be described.

Sulphur dioxide, sulphurous anhydride, SO₂ This substance is found in the emanations from volcanoes, and frequently in the air of towns, being derived in this latter case from the burning of coal and coal gas. It is the usual product of the

burning of sulphur, or the sulphides in air or in oxygen, and is generally obtained by such methods, when made on the large scale. For laboratory operations it is most conveniently obtained by deoxidizing sulphuric acid. The experiment usually requires strong acid and considerable heat. The substances which produce the best results are not soluble in the dilute acid. Copper, mercury, charcoal, silver, sulphur and other bodies may be used, but the first mentioned is the best for a small experiment.

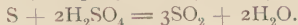
The gas comes off freely, and may be collected by downward displacement. The pneumatic trough can not be used, as the gas is very soluble in water. To show this properly, the end of the delivery tube may be dipped in some water in a beaker, when it will be found that most of the bubbles will be taken up by the liquid. This absorption must be carefully watched, or the water may be drawn back into the hot sulphuric acid.

The reaction in the preparation of sulphurous acid is somewhat difficult. When strongly positive metals, like zinc or magnesium, are put into acids, the usual result is the expulsion of hydrogen. With zinc and sulphuric acid we get, as before shown, $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + \text{H}_2$. With metals of less positive character, and therefore of lower affinity, a portion of the oxygen of the acid is removed, and water, sulphurous anhydride and a sulphate of the metal are formed. With copper we have



Mercury and silver also give similar effects. Several of the substances called non-metals, carbon and sulphur, for instance, also give a similar result, but no sulphates are formed, as these bodies can not form salts.

Carbon and sulphur give, respectively, the following reactions:



The carbon reaction is the most economical, but the admixed CO_2 is objectionable. Sulphur dioxide may also be made by heating an intimate mixture of manganese dioxide and powdered sulphur.



Properties. It is a colorless gas, of the well-known irritating odor of burning matches. It can be condensed to a colorless liquid by a cold 0°F . (-18°C .) A mixture of snow and salt answers quite well as a means of obtaining sufficient cold. The liquid so obtained is sulphur dioxide, SO_2 , and not sulphurous acid. It boils at 14°F . (-10°C .) and freezes at -105°F . (-76°C .) Sulphur dioxide passed into water forms sulphur-

ous acid, $\text{H}_2\text{O} + \text{SO}_2 = \text{H}_2\text{SO}_3$, which remains in solution, giving the liquid all the common properties of an acid; vegetable colors are first reddened, then bleached. The solution is also a powerful reducing agent, and is much used for that purpose.

Although theoretically we make a distinction between the anhydride and acid, yet, practically, we disregard this distinction, and for most experiments may use either the gas or the solution in water.

The principal properties of this body may be easily shown by the following experiments :

Dip a lighted taper in a jar in which the gas has been collected by displacement. The flame is immediately extinguished.

Suspend a delicately colored flower, somewhat moist, in a jar of the gas. The colors bleached by this agent are not entirely destroyed, and by exposure to the action of weak ammonia are often restored, with curious modifications.

Add some solution of sulphurous acid to the solutions of

(a) potassium bichromate with hydrochloric acid.

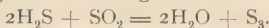
(b) potassium permanganate.

(a) will be changed to green chromic chloride.

(b) will be wholly decolorized. These effects are due to the reducing action of the sulphur dioxide, which thus becomes converted into sulphuric acid.

The anhydride, free acid and its salts are antiseptic agents, that is, prevent putrefaction and fermentation, especially the latter. For this reason, the vapors of burning sulphur are extensively used for fumigating ships and other places in which infectious diseases may exist. Wine and beer casks are also purified by sulphur; and the sulphites are added to fermented liquor to prevent further change. The acid is supposed to act by killing the minute living structures which are nearly always developed in the decomposing and fermenting substances.

Sulphur dioxide and sulphureted hydrogen decompose each other according to the following reaction :



The experiment may be easily performed by passing currents of the two gases into a small bottle.

Gen. Chem. Rel. The salts of sulphurous acid are called sulphites; monad metals, of course, form two compounds, acid and normal. Thus, potassium gives us

Acid potassium sulphite.



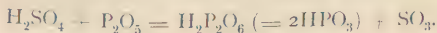
Potassium sulphite.



Dyad metals give but one sulphite. From calcium we have only CaSO_3 .

Tests. The odor and bleaching power of the free acid and the reducing properties of the sulphites are sufficient means of detection.

Sulphur Trioxide, Sulphuric Anhydride, SO_3 . This body is obtained by distilling Nordhausen sulphuric acid, (q. v.), and by the action of phosphoric anhydride upon common sulphuric acid. The latter reaction is a simple dehydration.



Several other methods of making it are known.

Properties. Sulphuric Anhydride is a soft, white, odorless solid, forming long, silky crystals like asbestos. Exposed to the air, it absorbs water rapidly and becomes converted into sulphuric acid. When dropped into water, the energy of combination is so great that a hissing noise is produced. The dry substance is entirely destitute of corrosive properties, and is now sent into commerce in sealed iron boxes for use in certain manufacturing operations. It is occasionally employed as a means of drying gases, and for special purposes in chemical research. It has a specific gravity of 1.95, melts at 65°F . (18.3°C .), and boils at 95°F . (35°C .).

Sulphuric Acid, Hydrogen Sulphate, Oil of Vitriol, H_2SO_4 . This substance was probably known to Geber in the 8th century, and was certainly known to Basil Valentine in the 15th century. It is now made in very large quantities in almost all parts of the civilized world, and it has been recently said, by a technologist, that the material prosperity of a country may be judged of by the extent of its sulphuric acid manufacture.

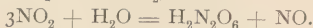
Sources. Sulphuric Acid occurs free in waters of volcanic and mining districts, and sometimes in the air of towns, being in the latter case derived from the oxidation of sulphurous acid. It also occurs in the saliva of certain animals. The compounds of sulphuric acid (sulphates) are of frequent occurrence. Calcium and barium sulphates are abundant minerals; sodium sulphate occurs in many natural waters.

Preparation. The original method of preparation was the distillation of the sulphates, especially the ferrous sulphate, FeSO_4 , which was produced by the oxidation of iron pyrites, FeS_2 . The process is still used for special purposes, as the acid so produced is more concentrated than the ordinary commercial article. The method by which the common acid is made depends upon the power of one of the oxides of nitrogen to act as carrier of oxygen from the air to sulphurous acid. The outline

of the practical process is as follows: Vapors of nitric and sulphurous acids are mixed with steam and air in a large leaden room, the floor of which is slightly inclined, and covered by a few inches of water. The sulphurous acid is derived either from burning of raw sulphur or the roasting of pyrites, the nitric acid from the reaction of sodium nitrate on sulphuric acid. The chemical changes are somewhat complicated, and are not wholly understood. The nitric acid changes some sulphurous acid to sulphuric, becoming itself converted into nitric oxide, NO , by this action. This NO takes oxygen from the air, and forms NO_2 , which oxidizes more sulphurous acid, and is thus again converted into NO , and again acted upon by the air. It will be seen that a small quantity of nitric acid will be sufficient to oxidize large quantities of sulphurous acid. Presence of large excess of water is essential to the reactions, hence steam or water in a fine spray, is thrown continuously into the room. The above description includes only the principal reactions; the following series of equations show more exactly the changes occurring:



The steam acts upon the NO_2 , thus



The NO takes O from the air and becomes NO_2 , which together with the nitric acid just produced oxidizes more sulphurous acid, reproducing the NO , and the changes go on anew.

The method may easily be shown experimentally:

Exp. A wide mouth quart bottle is provided, with a cork perforated for four tubes. To these tubes are attached a flask for generating sulphurous acid, a bottle for evolving nitric oxide, (q. v.) and a flask for furnishing steam; the fourth tube is left open to the air. Sulphurous acid and nitric oxide are produced, and allowed to flow into the bottle. They combine and produce a white crystalline solid, which quickly disappears when steam is admitted. Some water should also be poured down the open tube. The chemical changes are in part evident to the eye by the change of color which attends the conversion of red NO_2 into colorless NO , and *vice-versa*. After the experiment has lasted ten or fifteen minutes sufficient dilute acid will have collected in the bottle to respond to the tests given below.

In practice it is found that loss of NO is constantly occurring, and so it is necessary to renew the nitric acid occasionally. The liquid on the floor of the leaden room is drawn off from time to time, and concentrated by boiling in lead pans until it becomes strong enough to attack the lead. The further concentration is conducted in glass or platinum vessels.

Properties. The properties of sulphuric acid depend in part upon its purity, and degree of concentration.

Pure Sulphuric Acid is a colorless, oily liquid of a specific gravity of 1.848, boiling about 640° F. (338° C.) It is highly corrosive and poisonous. Exposed to the air it absorbs water in considerable amounts. When added to water it produces heat, often sufficient to make the water boil, and the dilution of any considerable quantity must be performed with care. A diminution of bulk occurs when the acid and water are mixed. So great is the affinity of sulphuric acid for water, that it will decompose many organic substances, extracting the hydrogen and oxygen and leaving the carbon. The carbon so liberated will diffuse through the acid, and give it a dark color.

Commercial Sulphuric Acid is usually more or less brown, or even black from the carbon set free from particles of dust, straw, etc., which accidentally fall into it. It always contains a small quantity of water—about one molecule to twelve of acid :



Its properties, boiling point, etc., are similar to those of the pure acid.

Nordhausen or Fuming Sulphuric Acid is the original oil of vitriol, so called because it was obtained by the distillation of green vitriol. It is substantially a solution of sulphur trioxide, SO_3 , in sulphuric acid. It is denser and even more corrosive than the common acid and unites with water with great energy. It is largely used for dissolving indigo and for a few other purposes. When heated, the sulphur trioxide distills off and the ordinary acid is left.

The properties of sulphuric acid are much modified by dilution; its corrosive and charring action may be entirely removed by adding much water. When such a dilute acid is boiled it steadily loses water until the original degree of concentration is nearly or quite restored. The following experiments show the more important properties :

Mix in a thin glass vessel, equal volumes of strong sulphuric acid and water, and place upon the surface of the liquid a small capsule, made out of tin or copper foil, and containing a small piece of phosphorus. The heat will be sufficient to ignite the phosphorus. If the quantities of water and acid be carefully measured the amount of condensation may be observed. It will be about one-fifth. The dilute acid so obtained is useful for many experiments and should be preserved.

The milkiess produced in this mixture is due to the precipitation of lead sulphate formed from the pans in which the acid is

concentrated. This body is soluble in the strong, but not in the dilute acid.

Place a few bits of straw, wood, or common organic matter in some strong sulphuric acid. In the course of a few minutes the acid will be quite discolored by the carbon set free. This discoloration of the acid does not interfere with its ordinary uses.

If any design be traced on white paper with the dilute acid obtained in the former experiment, and the paper then cautiously heated, the acid will slowly become more concentrated and will finally char the paper completely, but only at the parts which have been touched by the original liquid. This experiment is made use of for the detection of small quantities of the free acid.

The uses of sulphuric acid are very numerous. By its high affinity, it is capable of expelling many other acids from combination. Nitric acid is made from nitrates; acetic acid from acetates, by its agency. Its affinity for water makes it a useful drying agent especially for gases. Many organic bodies are peculiarly modified by treating with dilute sulphuric acid, but these changes will be best understood when described in the organic chemistry.

Gen. Chem. Rel. The salts of sulphuric acid are called sulphates. Monad metals give, of course, two sulphates, acid and normal. Sodium gives us

Acid sodium sulphate.



Sodium sulphate.



Dyad metals give but one sulphate. From barium we get BaSO_4 barium sulphate.

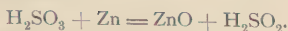
Most of the sulphates are soluble in water. Barium and lead sulphates are insoluble.

Tests. The concentrated acid is easily recognized by its oiliness and charring action on organic matter. The dilute acid may be made to produce this charring by evaporation, as described above.

The general test for either the acid or any of its salts is the addition of a solution of some barium compound (barium nitrate, chloride or acetate); a white precepsitate insoluble in water or dilute acids is at once formed, even if only a trace of the acid be present. The solution to be tested should be made acid, with hydrochloric acid, to prevent carbonates, phosphates, etc., being mistaken for sulphuric acid.

The commercial sulphuric acid contains several impurities. Of these the most important are arsenic and lead, which may be detected and removed by diluting the acid very much and adding sulphureted hydrogen.

Hyposulphurous Acid, H_2SO_2 is produced by dissolving zinc in sulphurous acid.



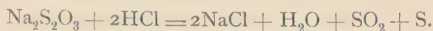
It is a powerful bleaching and reducing agent, and decomposes quickly.

Thiosulphuric Acid, $\text{H}_2\text{S}_2\text{O}_3$, commonly, but erroneously called hyposulphurous acid.

Sources. Calcium thiosulphate is found abundantly in the spent lime from gas-works, and in the refuse from soda ash manufacture. Small quantities of hyposulphites occur in many mineral waters.

Preparation. Sodium thiosulphate, or sodium hyposulphite as it is often called, is the only important salt. It is prepared by digesting sodium sulphite with sulphur, the reaction being a direct union of the two bodies, $\text{Na}_2\text{SO}_3 + \text{S} = \text{Na}_2\text{S}_2\text{O}_3$. It is also prepared by the action of sodium carbonate on calcium thiosulphate. The free acid has never been obtained. The properties and uses of the thiosulphates will be described under sodium. They are powerful reducing agents and can be used instead of sulphurous acid for that purpose.

Tests. When an acid is added to a thiosulphate, sulphur dioxide is evolved, and sulphur precepitated.



The precipitated sulphur renders the liquid turbid. This turbidity distinguishes the thiosulphates from the sulphites, as the latter do not throw down any sulphur.

A sesquioxide of sulphur, S_2O_3 , has been described but is not important.

The remainder of the sulphur acids have no particular interest and will not be described.

Carbon Disulphide, CS_2 . This body is precisely analogous to carbon dioxide. It is produced by passing vapor of sulphur over red-hot charcoal. It is a colorless liquid, which, when quite pure and in large quantity, has a rather pleasant odor, but when impure, and especially when diffused through the air in small quantity, is quite disagreeable. It has high dispersive power, that is, separates widely the various colors of the spectrum, and is often used in optical apparatus. It is very volatile and inflammable, and has high solvent powers, dissolving sulphur, phosphorus, and most oils and fats, and is much used for such purposes. Its vapor will take fire much below a red-heat.

Exp. Put a few drops of CS_2 at the bottom of a small glass vessel and allow them to evaporate. If a glass, heated to about

300° F. be introduced, the vapor inflames. The products of combustion are CO_2 and SO_2 . The analogy between CO_2 and CS_2 has been pointed out on page 59.

SELENIUM, Se.

ATOMIC WEIGHT 79.5. DENSITY 79.5.

Selenium was discovered by Berzelius in 1817.

Sources. It is found native and also in combination with various metals. It is rather rare. The method of preparation is not important.

Properties. The physical properties of selenium resemble those of sulphur. It shows several allotropic forms. The principal interest attaching to it is that its power of conducting electricity is affected by light. Several suggestions for apparatus for the electrical transmission of images have been based upon this property, and it is barely possible that by its means we may some day be able to *see* by telegraph.

Gen. Chem. Rel. The compounds of selenium are analogous to those of sulphur; we have

H_2Se	Hydrogen selenide.
SeO_2	Selenium dioxide.
H_2SeO_3	Selenous acid.
H_2SeO_4	Selenic “

The last is capable of dissolving gold.

TELLURIUM, Te.

ATOMIC WEIGHT 128. DENSITY 128.

Tellurium was discovered by Müller in 1782.

Sources. It is found native and also in union with bismuth, gold, etc. It is quite rare. The method of preparation is unimportant.

Properties. Tellurium has a metallic lustre and pinkish color. It fuses just below a red heat and at a temperature somewhat higher boils.

Its compounds are analogous to those of sulphur and selenium,

CHLORINE, Cl.

ATOMIC WEIGHT 35.46. DENSITY 35.46.

Chlorine was discovered by Scheele in 1774.

The name, derived from a Greek word meaning green, was given by Davy.

Sources. Chlorine is always found in combination. Its most abundant compound is common salt, NaCl, which is found in the animal, vegetable and mineral kingdoms. The chlorides of lead, silver and some other metals are also found as minerals. Of late years a considerable amount of potassium chloride has been obtained from the salt mines of Stassfurth, Germany. Hydrogen chloride is occasionally found in volcanic regions.

Preparations. Chlorine being extensively used in manufacturing chemistry, many processes for its preparation have been devised; nearly all of them depend upon the oxidation of some chloride. Hydrogen chloride, hydrochloric acid, is generally used. The following include the processes best suited for the laboratory:

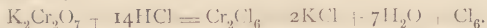
(a) By heating a mixture of manganese dioxide and hydrochloric acid.



(b) By heating a mixture of common salt, sulphuric acid and manganese dioxide,



(c) By heating a mixture of potassium bichromate and hydrochloric acid,



(d) By the action of dilute acids upon bleaching powder. This method is suitable for the preparation of very small amounts of chlorine, for use as a test, or for inhalation.

(e) Potassium chlorate mixed with hydrochloric acid, produces an impure chlorine which may also be used when only small amounts are required.

The process (b) is the most economical, (a) the most convenient for laboratory work.

Exp. Mix in a flask, about 1 ounce of black oxide of manganese with 4 ounces of hydrochloric acid; shake well, and heat the mixture gently. The chlorine is best collected by displacement. The operation should be conducted on the small scale and in a well ventilated place as the gas is very irritating. Narrow-mouthed, stoppered bottles of about 1 pt. capacity will answer

very well for receiving the gas. It cannot be collected over water or mercury.

Properties. Chlorine is a greenish-yellow gas, of a disagreeable and highly irritating odor. By a pressure of about sixty lbs. to the inch, it condenses to a greenish liquid which has never been frozen. The gas is about two and one-half times as heavy as air; one litre weighs 3,180.8 grms; water dissolves about three volumes, acquiring the color and odor of the gas; the solution, known as *chlorine water*, does not keep well.

The affinities of chlorine are very great. It combines with every element. It combines directly with most of the metals, decomposes water and bleaches and destroys many organic substances. Its affinity for hydrogen is increased by light.

Exp. A lighted taper put into a jar of chlorine continues to burn, but with a dark red flame, and the escape of clouds of smoke. The chlorine has combined with the hydrogen of the taper and not with carbon, hence the abundant liberation of the latter.

Exp. Powdered antimony dropped into chlorine takes fire at once and produces dense, irritating clouds of antimony chloride. Dutch leaf, a sort of brass, is also instantly burned up.

Exp. Paper dipped in oil of turpentine takes fire spontaneously in chlorine, producing a red flame and a dense cloud of smoke. This experiment requires pretty pure chlorine and good turpentine.

Exp. The bleaching power of chlorine is easily shown by pouring some solution of cochineal or litmus into a jar of the gas. The color is almost instantly removed. Pieces of calico may also be quickly bleached by placing them, in a wet state, in contact with the gas.

Exp. A small quantity of the gas as it comes from the evolution flask should be allowed to bubble through water. If a small quantity of this liquid be shaken in a bottle containing hydrogen sulphide, the odor of this gas will disappear.

The last two experiments show the bleaching and disinfecting applications of chlorine. It is, however, rarely used in form of gas, on account of the obvious inconvenience. It is usually employed in the form of bleaching powder, made by passing the gas into slacked lime. The body so produced is more fully described among the calcium salts. It is easily decomposed by dilute acids, yielding its chlorine; even the carbonic acid of the air will act upon it. It is generally employed in solution and is often incorrectly called chloride of lime.

Chlorine does not bleach unless moist, and it is believed that

a decomposition of water first occurs, and that the oxygen thus set free is the active agent.



The powerful action of chlorine upon metals is utilized occasionally for the decomposition of certain ores.

Gen. Chem. Rel. In the chlorides the chlorine is electro-negative and monatomic. It is capable, however, of assuming electro-positive relations, and higher atomicities; and in its compounds with oxygen shows both these changes. It is also capable of replacing hydrogen, atom for atom, thus giving rise to an important and extensive series of *substitution* compounds, which are best considered in connection with the organic chemistry. The chlorides are generally less numerous than the corresponding oxides; for instance, we know two oxides of hydrogen, but only one chloride.

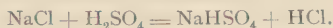
Hydrochloric Acid, Muriatic Acid, Spirit of Salt, HCl. This substance was long known in an impure form. Basil Valentine (15th century), described the pure acid, and Davy, in 1810 showed its composition.

Sources. Hydrochloric acid occurs in the gases evolved from volcanoes, and in solution in the waters of some mountain streams of South America.

Preparation. The acid may be formed by the direct union of its elements, but this method has only theoretical interest. The practical process is the action of common salt and sulphuric acid, according to the following reaction:



This reaction requires a high temperature. In ordinary experiments and on the small scale the reaction is



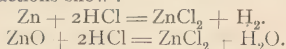
By which, from a given amount of sulphuric acid, only half the quantity of hydrochloric acid is obtained. This form of the process is therefore costly.

Exp. One ounce of common table salt is mixed with twice its weight of strong sulphuric acid, in a flask provided with a funnel, and delivery tube in the ordinary manner. The hydrochloric acid comes off freely, and may be collected by downward displacement. Toward the latter part of the operation a gentle heat may be used. A portion of the gas should be passed into water.

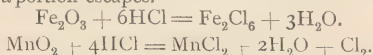
Hydrochloric acid is a colorless gas of a strong pungent odor, and poisonous to animals and plants. Its density is 1.8181; 1 litre weighs 1.63 grms. It may be liquified by a pressure of about six hundred pounds to the inch. It does not burn nor sup-

port ordinary combustion, but many metals burn in it, forming chlorides. Its most important property is its solubility in water, which at low temperatures will absorb nearly 500 volumes, producing a strongly acid solution, which is the common hydrochloric or muriatic acid. The strongest form usually sold contains 43 per cent. by weight, of the gas. When pure the solution is a colorless, fuming, strongly acid liquid, but the commercial forms are usually yellow from the presence of iron.

Gen. Chem. Rel. Hydrochloric acid is rather weaker than sulphuric or nitric acid, but is used largely as a solvent. It generally acts by forming chlorides. When metals are dissolved hydrogen escapes; when oxides are dissolved water is formed, as the following reactions show:



With oxides more rich in oxygen, ($\text{MnO}_2, \text{Fe}_2\text{O}_3, \text{CrO}_3$), the action is dependent upon the temperature and upon the atomicity of the metal. Sometimes the whole of the chlorine is retained, in other cases a portion escapes.



Tests. Hydrochloric acid produces with ammonia, white fumes of NH_4Cl . In common with all the chlorides it produces, with silver nitrate, a white curdy precipitate of silver chloride, soluble in ammonia.

A mixture of about three parts nitric with five parts hydrochloric acid has been long used under the names, aqua regia and nitromuriatic acid. It dissolves gold and platinum and owes its efficacy in part to the free chlorine which is formed by the oxidizing action of the nitric acid upon the hydrogen of the muriatic.

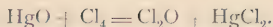
Compounds of Chlorine and Oxygen. Three of these are known in the free state:

Cl_2O	Hypochlorous anhydride.
Cl_2O_3	Chlorous anhydride.
ClO_2	Chlorine peroxide.

Several others are known only in union with the metals, and in these we have a well marked and important series:

HClO	Hypochlorous acid.
HClO_2	Chlorous “
HClO_3	Chloric “
HClO_4	Perchloric “

Hypochlorous anhydride, Cl_2O , is formed when chlorine is passed over mercuric oxide,



The combination of oxygen and chlorine will not take place when the two elements are mixed under ordinary conditions. The action in the present case is due to the fact that the oxygen is just liberated from combination, and its affinity is temporarily increased. Quite a number of instances are known in chemistry in which bodies can be made to combine by bringing them in contact at the moment they have separated from other compounds, in the nascent state, as it is often called, that is just born. Hypochlorous anhydride is absorbed by water and is supposed to form hypochlorous acid, HClO , which has been used as a bleaching agent, especially for removing ink stains.

Exp. Shake a few grains of finely powdered mercuric oxide with some chlorine water. The odor of chlorine will be replaced by that of hypochlorous acid, and the liquid will easily remove stains of writing without seriously injuring the paper. Other hypochlorites may be formed by the action of chlorine upon metallic oxides or hydrates at a *low* temperature, (see calcium hypochlorite.)

Chloric Acid, HClO_3 . If the action of chlorine upon metallic oxides or hydrates be at a temperature of over 60°F. (15.5°C.), *chlorates* will be produced, according to the following reaction:



The chlorate and chloride are separated by difference of solubility in water. Chloric acid may be obtained from the chlorates by stronger acids but it has no practical value.

The chlorates are useful for the large amount of oxygen which they contain and which they give up easily when heated. Potassium chlorate is the great source of pure oxygen as mentioned on page 32. It is used largely in fireworks. Perchloric acid is obtained by heating dilute chloric acid. The perchlorates resemble the chlorates.

The other compounds of chlorine and oxygen have no practical importance.

Compounds of Chlorine with Sulphur and Carbon. These have considerable theoretical and some practical interest, but will be only briefly mentioned. With sulphur we have

S_2Cl_2	Sulphur Chloride.
SCl_2	“ Dichloride.
SCl_4	“ Tetrachloride, (doubtful.)

The first two are liquids. S_2Cl_2 is used as an agent in vulcanizing rubber.

With carbon we have at least four compounds :

C_2Cl_2	Carbon Monochloride.
C_2Cl_3	“ Dichloride.
C_2Cl_6	“ Trichloride.
CCl_4	“ Tetrachloride.

They cannot be prepared by the direct union of their elements, but are mostly the result of the successive substitution of chlorine for hydrogen. Marsh gas, CH_4 , for instance, yields, by such action, the following compounds :

CH_3Cl	Monochlorinated marsh gas.
CH_2Cl_2	Dichlorinated “ “
$CHCl_3$	Trichlorinated “ “
CCl_4	Tetrachlorinated “ “

The third body, $CHCl_3$, is chloroform; the fourth is carbon tetrachloride, the most important of the carbon chlorides. It is a colorless, volatile liquid, which acts as a powerful anæsthetic.

Chlorine combines with nitrogen to form a body called nitrogen chloride, of which the composition is somewhat uncertain. It is an oily liquid which decomposes very easily and with a violent explosion.

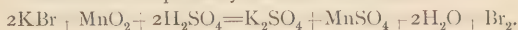
BROMINE, Br.

ATOMIC WEIGHT 80. DENSITY 80.

Sources. Occurs in sea water and sea plants, in brine springs, and in a few minerals. Its compounds are generally associated with those of chlorine. Considerable quantities have of late been obtained from brine springs in western Pennsylvania. Bromine was discovered by Balard in 1826. The name is derived from a Greek word meaning “bad smell.”

Preparation. Bromine is prepared by processes analogous to those of chlorine, acting upon bromides by means of oxidizing agents, such as a mixture of sulphuric acid and manganese dioxide. It may also be directly expelled by the superior affinity of chlorine.

Exp. Dissolve some potassium bromide in a little water, add a pinch of manganese dioxide and a few drops of sulphuric acid. Heat gently, and dark red vapors of bromine will soon be evolved. The reaction is precisely like that for chlorine :



Exp. To a solution of potassium bromide in water, add a few

drops of chlorine water, and shake. The bromine will be set free and color the water yellowish. If some ether be now poured in and shaken for a few moments, the bromine will leave the water and dissolve in the ether, which will form a red layer on the surface of the water. The reaction is



The free bromine is usually dissolved in potassium hydrate, by which a mixture of potassium bromate and bromide is produced. The reaction is similar to that which occurs in the preparation of potassium chlorate, (page 73.) This mixture being heated, the bromate loses oxygen and forms bromide.

A considerable amount of bromine is now sent into commerce in the free state.

Properties. Bromine is a dark red liquid, which at ordinary temperatures evolves red vapors of an excessively irritating and disagreeable odor. The liquid is three times as heavy as water, and boils at 145°F . (61°C .), and freezes at -12°F . (-25°C .) It is soluble in water, and is often conveniently used in that form. Its chemical properties are similar to those of chlorine, but are not so energetic. It bleaches vegetable colors, and by decomposing water acts as an oxidizing agent. A number of its compounds are used in medicine.

Exp. The affinity of bromine may be easily shown by placing a few drops on a small piece of phosphorus, which will at once be ignited.

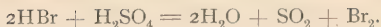
Gen. Chem. Rel. The chemical relations of bromine are almost exactly those of chlorine. It combines with the metals, forming bromides, of which those of hydrogen, potassium and ammonium are the most important. In these compounds the bromine is a monad. It also forms oxygen compounds analogous to those of chlorine, but they have very little interest.

Tests. Bromides may be recognized by the production of the red vapor of bromine by the action of free chlorine or of a mixture of manganese dioxide and sulphuric acid. With solutions of bromides, silver nitrate gives a light yellow precipitate of silver bromide which is slightly soluble in ammonia.

Hydrogen Bromide, Hydrobromic Acid, HBr. This substance cannot be conveniently prepared by the action of sulphuric acid upon a bromide, which would be the theoretical method, because the sulphuric acid is decomposed by the hydrogen, and instead of the reaction



we have the HBr acting on another portion of the sulphuric acid, and giving



Hydrobromic acid is obtained by using a mixture of phosphorus, powdered glass and bromine, or by the action of phosphoric acid upon a bromide. It resembles hydrochloric acid in its properties, and has been used in medicine.

Bromic Acid, HBrO_3 , and **Hypobromous Acid**, HBrO , are also known. They closely resemble the corresponding chlorine compounds.

IODINE, I.

ATOMIC WEIGHT 127. DENSITY 127.

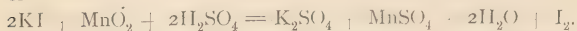
Sources. Occurs in association with bromine and chlorine in sea-water and sea-plants. The latter are burned, and the ashes, called *kelp*, contain various iodides.

Iodine was discovered by Courtois in 1811. The name means violet colored, and refers to the vapor.

Preparation. Iodine is prepared by processes exactly similar to those for bromine, either by the action of chlorine, or of the mixture of manganese dioxide and sulphuric acid. The reactions are



or



The reactions may be shown by substituting an iodide for a bromide in the experiments described on page 74.

Properties. Iodine forms bluish-black crystalline masses, with a metallic lustre. It evaporates slowly, at ordinary temperatures, melts at 225° F. (175° C.) The vapor has a deep violet color and a peculiar odor, somewhat like that of chlorine, but not so irritating. The solid dissolves slightly in water — much more freely in water containing potassium iodide, and in alcohol, ether, and carbon bisulphide. It has some bleaching and oxidizing powers. One of its important properties is the power of producing a blue color with starch. For this action the iodine must be the free state; the iodides give no color.

Exp. Prepare some starch solution by boiling common starch with enough water to make a thin liquid. To one portion of this liquid add a few grains of potassium iodide; to another portion add a drop or two of solution of iodine. The first solution will remain colorless; the second will become deep blue. Boiling

will remove the color. Dip slips of paper into the colorless solution and expose them to either the vapors of nitric acid, chlorine or ozone; a blue color will be immediately produced because the iodine is set free. Papers prepared in this manner are, therefore, used as tests for the vapors just mentioned.

Exp. The affinity of iodine is shown by adding a few grains to a small piece of phosphorus. Combustion occurs in a few seconds, and if a large bell jar be set down over the mass, a quantity of iodine vapor will be collected in it.

Gen. Chem. Rel. The chemical relations of iodine are substantially the same as those of chlorine and bromine.

Tests. The violet vapor and the blue color with starch are characteristic. For iodides, silver nitrate is used; it gives a yellow precipitate, insoluble in ammonia.

Hydriodic Acid, Hydrogen Iodide, HI. This is prepared by methods similar to those used for hydrogen bromide, which body it closely resembles. It has been used in medicine.

Two compounds with oxygen are known, of which iodic acid, HIO_3 , has some little importance, from being used as a test for morphia, which produces with it a brown color.

By the action of strong ammonia upon powdered iodine, a brownish substance is produced, which was supposed to be nitrogen iodide, but probably contains hydrogen. It is easily handled while wet, but when perfectly dry explodes, with a loud report, on the slightest touch.

Exp. Put a few grains of iodine into a watch glass, and pour on enough strong ammonia to cover the mass. Crush the iodine with a glass rod, and allow the mixture to remain quiet for five or ten minutes; pour off the ammonia, wash with water, and put the dark brown powder on small pieces of filter or blotting paper, and allow to dry, which takes from 30 to 60 minutes. The slightest touch, even a breath of air, will be sufficient to cause the explosion.

FLUORINE, F.

ATOMIC WEIGHT 19 DENSITY, UNDETERMINED.

Soures. Tolerably abundant as Fluor Spar, CaF_2 , and Cryolite, 6NaF , Al_2F_6 , and some rarer minerals. It exists in the stems of grasses, and in bones and teeth.

Preparation and Properties. Fluorine has never been satisfactorily prepared. Its high affinities and its special power

of acting on glass and on metals render it difficult to experiment with it. It has been described as a yellow gas, and also as colorless.

Gen. Chem. Rel. It combines with every known element except oxygen, and bears a close resemblance to chlorine in most of its relations. It is especially remarkable for its affinity for silicon.

Hydrogen Fluoride. Hydrofluoric Acid. HF . This body is easily prepared by acting on calcium fluoride, CaF_2 , with sulphuric acid. The operation must be performed in vessels of lead or platinum. The pure HF is a gas, but it is commonly seen as a strong solution. It acts powerfully on most metals, and on siliceous materials. It is used for etching designs on glass.

Exp. Prepare a glass plate, flowing over it some wax or paraffine, and cutting away the coating in any design. Invert the coated plate over a leaden dish in which has just been placed some strong sulphuric acid, and enough powdered calcium fluoride to make a thick paste. The action may be assisted by a gentle heat, but care must be taken not to melt the wax. After ten or fifteen minutes the uncovered parts of the glass will be found corroded. The wax can be gotten off by warming the plate.

Strong solution of hydrogen fluoride is now sold in gutta percha bottles, upon which it has no action.

Tests. Fluorine compounds are recognized by their power of producing hydrogen fluoride when heated with sulphuric acid. The hydrogen fluoride is easily detected by its action on glass.

PHOSPHORUS, P.

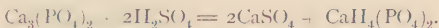
ATOMIC WEIGHT, 31. DENSITY, 62.

Sources. Occurs principally as calcium phosphate, which exists in bones and teeth, in many minerals, and in soils. Various phosphates also exist in the fluids of the animal body. Discovered by Brandt in 1669. The name means "carrier of light."

Preparation. Phosphorus is prepared from bones which contain from one-third to two-thirds their weight of calcium phosphate. The bones are deprived of their animal matter, and the *bone-ash* thus left is treated with sulphuric acid, by which a soluble acid calcium phosphate is formed, and much calcium sulphate is deposited. The liquid is then concentrated, mixed

with charcoal and sand and heated. Calcium silicate is produced, the charcoal takes the oxygen, and the phosphorus distills over. The reactions are complicated; the following is to be considered as merely an indication of the change and not the exact proportion.

Bone ash.



Acid cal. phosph. Sand.



The CaSiO_3 represents the calcium silicate but is not the correct formula.

Other methods of manufacture are known, but need not be described. The distilled phosphorus is collected in water and cast in sticks.

Properties. Phosphorus when freshly prepared is a colorless, almost transparent solid, soft as wax; when kept for sometime, especially in the light, it becomes brownish, opaque and harder. It takes fire easily and is usually kept under water. It burns with a bright flame producing white clouds of phosphoric anhydride, P_2O_5 . Exposed to the air at low temperature it can still undergo a slow combustion, producing P_2O_3 ; it is then luminous in the dark. It is insoluble in water but dissolves in oils and in carbon disulphide. It is extremely poisonous, death having occurred from much less than $\frac{1}{4}$ grain. Phosphorus melts at $111^\circ \text{F. } 43^\circ \text{C.}$, and boils at $550^\circ \text{F. } (285^\circ \text{C.})$, producing a vapor which according to analogy should be 31 times as heavy as hydrogen, but is really 62 times as heavy. By keeping it at a temperature of $450^\circ \text{F. } (230^\circ \text{C.})$, for days, in a closed vessel, phosphorus is converted into the

Amorphous or Red Variety, an allotropic form in which most of its properties are changed. In this form it is red, insoluble in carbon disulphide, difficult to burn, non-poisonous, and shows many other minor differences. Its composition is the same. This change is also produced by adding a small quantity of iodine to common phosphorus.

The uses of the element in matches and as a medicinal substance are well known.

In all experiments with it, great care must be taken, as it is easily inflamed and produces one of the most severe forms of burns known. It should be handled with a pair of forceps and cut or divided only under water.

Gen. Chem. Rel. Phosphorus belongs to the nitrogen group, and acts as a triad or pentad; its affinities in the free state are, however, very high. It forms compounds with many elements,

but those with metals are not of much interest, except with hydrogen and iron. It is a powerful reducing agent. Its important properties are shown in the experiments given on pages 33, 40, 75, and 77.

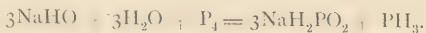
Tests. Phosphorus in the free state is easily recognized. In very minute quantity it is detected by its luminosity when distilled in a dark room.

The formation of the amorphous variety may be noticed by adding a few grains of iodine to a small fragment of common phosphorus.

Hydrogen Phosphide, PH_3 , Phosphine. This body is formed under conditions analogous to those which produce ammonia, that is, when its elements are brought together in the nascent state, (page 73.) When a solution of caustic alkali is boiled with phosphorus, water is decomposed and the hydrogen phosphide is formed.

Exp. Some small pieces of phosphorus are put into a small retort, which is then *nearly filled* with a strong solution of caustic soda, and arranged so that the end of the neck can be quickly put below the surface of some water in a basin. The retort should then be heated by a spirit lamp, and when the mixture boils the hydrogen phosphide will come off. The end of the retort should then be put below the water. The gas bubbles through the water and burns on reaching the air. The experiment requires care and should be performed only by those who have some acquaintance with such manipulations. When the action is over, it is best to simply leave the apparatus to itself, as a slight explosion generally occurs.

The reaction is



NaH_2PO_2 is sodium hypophosphite. Hydrogen phosphide is a colorless gas of a disagreeable odor. As ordinarily made it is spontaneously inflammable, but this property is due to the presence of a small quantity of the vapor of a liquid phosphide, PH_2 . If this latter be removed by passing the fresh gas through a tube placed in a freezing apparatus, the power of inflaming spontaneously is lost. A solid phosphide, apparently P_2H is also known.

The inflammable gas can also be prepared by the action of water on calcium phosphide.

Hydrogen phosphide has no alkaline properties, but forms many compounds analogous to those formed by ammonia.

Compounds of Phosphorus with Oxygen. Only two compounds are definitely known. These are :

P_2O_3 Phosphorous anhydride.

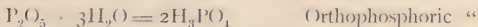
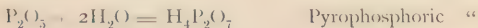
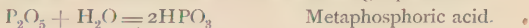
P_2O_5 Phosphoric anhydride.

Phosphorous Anhydride, P_2O_3 , is produced by the slow oxidation of phosphorus. When this takes place in ordinary air, water is absorbed and phosphorous acid, H_3PO_3 , is formed. This is a powerful reducing agent but is not an important body.

Phosphoric Anhydride, P_2O_5 . This is the product of the active combustion of phosphorus. It is easily produced by burning phosphorus in the air. It forms white, snow-like fumes which rapidly absorb water.

Exp. Ignite a small piece of phosphorus on a common plate, and cover it with a jar. The white clouds of P_2O_5 soon begin to settle on the plate, and when the combustion is finished, throw a few drops of water on the white deposit; a hissing sound will be heard, due to the energetic union of the anhydride with the water.

Phosphoric anhydride is a white, snow-like solid, having a very high affinity for water. It is capable of uniting with water in at least three proportions, forming different bodies; very few anhydrides show this power, and consequently, the chemistry of these phosphoric acids is more than ordinarily complicated. When the anhydride is mixed with water as in the above experiment, it combines with one molecule of water and forms HPO_3 . If, however, the acid be obtained from any of the phosphates found in nature, it has the formula H_3PO_4 . It can be produced from the anhydride by heating with three molecules of water. An intermediate acid, $H_4P_2O_7$, is also known. The relation between these different forms is shown in the following equations:



The last acid is the one that yields all the natural phosphates; the syllable ortho means "regular" and signifies that this is the regular form of the acid. The syllable pyro means "fire," and signifies that the second acid, $H_4P_2O_7$, is obtained by the action of heat upon some of the natural phosphates. By the further action of heat the metaphosphoric is obtained.

Properties and Gen. Chem. Rel. of the phosphoric acids. Metaphosphoric and pyrophosphoric acids are of comparative unimportance. They are artificial products of the laboratory. The first mentioned is distinguished by a power of coagulating albumin. It is to be particularly noticed that although the three phosphoric acids differ in oxygen, the termination "ic" is not changed. This is because they are all formed from the same

anhydride; the difference in oxygen is due to the amount of water. The number of salts formed by each acid is in proportion to the number of molecules of water which it has taken up.

Metaphosphoric Acid, produced by adding *one* molecule of water, gives *one* series of salts.

NaPO_3	Sodium metaphosphate.
$\text{Ca}(\text{PO}_3)_2$	Calcium “

Pyrophosphoric Acid is produced by adding *two* molecules of water, gives *two* series of salts.

$\text{Na}_2\text{H}_2\text{P}_2\text{O}_7$	Acid sodium pyrophosphate.
$\text{Na}_4\text{P}_2\text{O}_7$	Sodium “

Orthophosphoric Acid produced by adding *three* molecules of water, gives *three* series of salts.

NaH_2PO_4	Acid sodium orthophosphate.
Na_2HPO_4	“ “ “
Na_3PO_4	Sodium orthophosphate.

Further consideration of these phosphates will be found in the introduction to the chemistry of the metals.

Tests. The detection of metaphosphoric and pyrophosphoric is not often required, but orthophosphoric acid is a body very often encountered in analysis. Silver nitrate produces with it a yellow precipitate, soluble in ammonia. A mixture of magnesium sulphate, ammonia and ammonium chloride gives a white precipitate soluble in acids.

A solution of ammonium molybdate in nitric acid gives a bright yellow precipitate insoluble in acid. This is a very delicate test.

The compounds of phosphorus with other elements are of limited importance. With sulphur it gives a variety of combinations. The action between the two elements may be shown by igniting a piece of phosphorus, placed in the midst of a mass of finely powdered sulphur. A quick combustion, producing a large flame, at once occurs. The phosphorus sulphide first formed is probably quickly consumed.

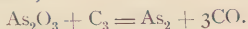
Two phosphorus chlorides are known, PCl_3 and PCl_5 . They have been much used in researches in organic chemistry, Compounds with bromine and iodine are also known,

ARSENIC, As.

ATOMIC WEIGHT, 75. DENSITY, 150.

Sources. Occurs in the free state, and as sulphide, also in combination with various metals, especially nickel, cobalt and iron. It is rather abundant and exists in small amounts in many minerals.

Preparation. Arsenic is prepared by deoxidizing arsenous anhydride by charcoal,



It was first prepared by Brandt, in 1733.

Properties. Arsenic, when freshly prepared, is a steel-grey brittle mass with a decidedly metallic lustre. It tarnishes somewhat in the air, and passes into vapor at about 356°F. (180°C.), without fusing. The vapor has a density twice as great as analogy would require. Heated in contact with air, it oxidizes to arsenous anhydride, and develops a garlicky odor. Arsenic is not dissolved by any simple solvent, and this fact, together with its lustre, led to its being formerly classed with the metals.

Gen. Chem. Rel. Arsenic is very similar to nitrogen and phosphorus in its chemical relations.

Tests. The tests for arsenic are given in connection with the chemistry of arsenous anhydride.

Arseneted Hydrogen, Arsine, AsH_3 . This body is analogous to ammonia, and like it, is produced when its elements are brought together in the nascent state. Its formation is one of the most delicate tests for arsenic. The usual method of preparation is to liberate hydrogen in a solution of arsenous anhydride. It is a combustible gas of disagreeable odor and excessively poisonous.

The important points in the chemistry of this body are given in connection with Marsh's test.

A solid arsenous hydride of uncertain composition is known.

Compounds of Arsenic with Oxygen. These are two:

As_2O_3 Arsenous oxide or anhydride.

As_2O_5 Arsenic " "

Arsenous Anhydride, Arsenous Oxide, White Arsenic, As_2O_3 . This is the substance generally called arsenic. It is obtained by roasting arsenical ores in a current of air, and presents itself in commerce in two varieties:

(a) The vitreous form, transparent and colorless at first, but afterwards becoming yellowish and porcelain-like.

(b) A pulverulent form, which is distinctly crystalline.

Properties. Arsenous anhydride is a white solid, odorless and tasteless, and dissolving with difficulty, and only in small amounts in cold water; the solution is feebly acid and is supposed to contain arsenous acid H_3AsO_3 . Hot water is a more active solvent, but the amount dissolved is dependent on many conditions, and variously stated by different observers. As a rough statement it may be said that a fluid ounce of cold water will dissolve about one grain, and the same amount of water if kept for one hour at the boiling point will take up about forty-five grains. In acid and alkaline solutions it dissolves much more readily. Heated to 380°F . (193°C .), the solid passes into vapor without fusing, and if allowed to condense, produces brilliant transparent crystals. It is intensely poisonous in all its forms, two or three grains being a fatal dose. Arsenous anhydride is used in making opaque white glass; in various solutions for preserving animal skins, and in the manufacture of colors. Its frequent occurrence and poisonous qualities have made its properties and tests of great importance. The following is a brief summary of the methods used.

1. *Reduction test.* This depends on the conversion of the oxide into the elementary arsenic. A small quantity of powdered white arsenic is mixed with some charcoal or some dried potassium ferrocyanide and heated in a narrow glass tube. The elementary arsenic is set free, rises in vapor, and condenses on a cooler portion of the tube, as a dark steel-grey, but rather lustrous layer called the *arsenical mirror*. If this deposit be heated it may be driven further along the tube, and will finally oxidize and produce the garlicky odor.

2. *Sublimation test.* Arsenous oxide heated alone passes quickly into vapor, and by allowing this vapor to condense upon a *slightly warmed* part of the tube, fine crystals are formed. Under the microscope these crystals are seen to be octahedral, that is, consist of eight triangular faces, though they are rarely completely formed. Very minute quantities of arsenic can be recognized by this test.

3. *Reinsch's test.* This is the most valuable test because it can be applied to impure mixtures, as the contents of a stomach. A small quantity of water is put into a wide test-tube or porcelain basin, some hydrochloric acid is added, a piece of clean copper is put in and the water brought to boiling. A few drops of a solution of arsenic are now added and in a few seconds a rather dull steel-colored deposit of copper arsenide forms on the copper. When this deposit has become rather dense, the copper is taken out, dried with filter paper, rolled up into small bulk and placed

in the end of a small glass tube. Heat being applied, the arsenical deposit is oxidized and volatilized, forming octahedral crystals of arsenous anhydride.

4. *Marsh's test.* This depends on the power of nascent hydrogen to form AsH_3 . The hydrogen is obtained either by the action of sulphuric acid upon zinc or magnesium, of sodium amalgam on water or by a current of electricity. The test is very delicate but requires great skill in its manipulation. The simplest method of performing it consists in mixing zinc and dilute sulphuric acid in a gas bottle, allowing the hydrogen to flow for some time, and then introducing a small quantity of arsenical solution. The arsine, AsH_3 , begins at once to come off, the flame of the hydrogen becomes livid and gives off fumes of arsenic. If a cold porcelain plate be held in the flame an arsenical soot will be deposited as a brown shining stain. If the tube which is conducting the current be heated, the gas will be decomposed and a similar stain formed within the tube. The stains may be identified as arsenic by the fact that they are :

(a) easily volatile.

(b) soluble in a solution of bleaching powder.

(c) capable of producing octahedral crystals of As_2O_3 .

Three tests, known as the liquid tests, are applicable only to pure solutions of arsenous anhydride. They are :

1. Hydrogen sulphide produces a lemon yellow precipitate of arsenous sulphide, As_2S_3 ,



A few drops of hydrochloric acid facilitate the action.

2. Silver nitrate, made alkaline by ammonia, gives a yellow precipitate of silver arsenite.

3. Copper sulphate, made alkaline by ammonia, gives a green precipitate of copper arsenite.

Arsenic Anhydride, As_2O_3 . This is produced by oxidizing arsenous anhydride with nitric acid. It forms, with water, arsenic acid H_3AsO_4 , which is used as an oxidizing agent in the manufacture of aniline colors. This use has been supposed to account for the cases of skin irritation which have been occasionally observed to follow the wearing of fabrics dyed with these colors, but it is very doubtful if any arsenic remains in the manufactured article.

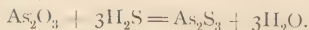
Arsenic acid forms salts called arsenates. It is usually tested by first converting it into arsenous acid. Three forms of arsenic acid are known, corresponding to the three forms of phosphoric acid.

Compounds of Arsenic and Sulphur. Three of these are known.

As_2S_2	Arsenous disulphide, Realgar.
As_2S_3	Arsenous sulphide, Orpiment.
As_2S_5	Arsenic “

Realgar is a brick-red solid, easily volatile. It is found as a mineral and may also be produced artificially. It has little practical importance.

Orpiment, King's yellow, is found as a mineral, and is easily produced artificially by the action of hydrogen sulphide upon arsenous anhydride,



It is a bright yellow solid, fusible and volatile, soluble in alkalis but insoluble in water and dilute acids. It is often obtained in the process of testing for arsenic, and in the arts is used as a pigment.

Arsenic Sulphide is unimportant.

Arsenic forms chlorides bromides and iodides, but they need not be described.

ANTIMONY, Sb.

ATOMIC WEIGHT, 122. DENSITY, 244, (PROBABLY.)

Sources. Occurs sometimes in the free state, but generally as sulphide, Sb_2S_3 . Discovered by Basil Valentine, in the 15th century. It is also called *stibium*.

Preparation. By melting the roasted antimony sulphide with charcoal and sodium carbonate. The reaction is analogous to that occurring in the preparation of arsenic.

Properties. Antimony is decidedly metallic in appearance, is still classed among the metals by many chemists, and may be regarded as a connecting link between the two classes of elements. It is bluish-white, brittle, generally highly crystalline and of brilliant lustre. It fuses at 842°F. (450°C.), and volatilizes at a red heat. On cooling from the melted condition it expands somewhat, and some of its compounds with metals retain this property, for which reason the element is much used in making alloys, like type-metal, which take sharp casts. Like arsenic, it is not soluble in any simple solvent.

Exp. Heat a piece of antimony, about the size of a large shot,

on a piece of charcoal, by means of the blowpipe, and when the metal is in perfect fusion drop it from a height of several feet upon a large sheet of paper. The globule will splash and throw little globules in all directions, each of which leaves a train of antimony oxide.

Gen. Chem. Rel. The chemical relations of antimony are very much like those of arsenic, phosphorus and nitrogen, and in spite of its metallic appearance it is now placed with the non-metals, but unlike them it possesses a limited power of forming salts.

Tests. Antimony is detected by tests similar to those of arsenic. The distinctive differences are :

1. The sublimate of free antimony cannot be obtained by the reduction test, unless a very high temperature be used.
2. The antimony oxide, cannot be volatilized except by a high heat, and does not usually form octahedral crystals, but these have been obtained under certain conditions.
3. The copper slip in Reinsch's test becomes covered with a bluish or violet deposit, which gives a sublimate only with great difficulty.
4. In Marsh's test, a much darker spot is obtained ; it is volatilized with difficulty and not dissolved by a solution of bleaching powder.

5. The liquid tests give no result except with hydrogen sulphide, which produces an *orange-red* precipitate.

Stibine, Antimoneted Hydrogen, SbH_3 , resembles the corresponding arsenic compound and is produced under similar conditions. It has not been obtained pure. The distinction between it and arsine is given above.

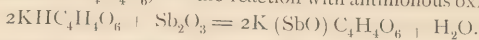
Compounds of Antimony with Oxygen. These are :

Sb_2O_3	Antimonous oxide, or anhydride.
Sb_2O_5	Antimonic " "

An intermediate oxide, Sb_2O_4 , is known, but it is generally regarded as a compound of the other two.

Antimonous Oxide, Sb_2O_3 . This is found as a mineral and is also readily prepared by burning antimony in the air. It is like As_2O_3 in many of its chemical relations, but is insoluble in water, less volatile, and shows some power of combining with acids to form salts, thus resembling the true metallic oxides. When boiled with a solution of cream of tartar, (acid potassium tartrate), antimonous oxide loses one atom of oxygen, and dissolves, forming *tartar emetic*, potassium antimony tartrate. This compound is the most familiar preparation of antimony, as unlike

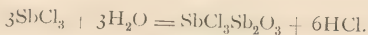
the antimony salts generally, it is soluble in water without decomposition. The composition is exceptional; acid potassium tartrate is $\text{KHIC}_4\text{H}_4\text{O}_6$, and the reaction with antimonous oxide is,



The SbO here acts like a metal, and replaces the hydrogen. Antimonous oxide also combines with strong bases, forming salts, called antimonites.

Antimonic Oxide, Sb_2O_5 , forms two acids corresponding to the meta- and pyrophosphoric acids. Unfortunately, the names of the antimony acids have been misplaced; HSbO_3 has been called antimonic acid, and $\text{H}_4\text{Sb}_2\text{O}_7$ metantimonic. This latter name should be given to HSbO_3 ; and $\text{H}_4\text{Sb}_2\text{O}_7$ should be called pyrantimonic acid. Pyrantimonic acid is remarkable for forming the only sodium compound insoluble in water. The orthantimonic acid, H_3SbO_4 has not yet been obtained.

Antimony forms compounds with chlorine, bromine and iodine analogous to those of phosphorus and arsenic. They are mostly decomposed when mixed with large quantities of water, yielding at first an impure, finally a pure oxide. With antimonous chloride we have



The oxychloride $\text{SbCl}_3\text{Sb}_2\text{O}_3$ becomes finally converted into pure antimonous oxide.

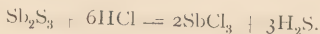
Antimony Sulphides. Two are known:

Sb_2S_3 Antimonous sulphide.

Sb_2S_5 Antimonic sulphide.

Antimonous Sulphide is the principal ore of antimony.

It is found as a shining, grey, crystalline mass, fusible and easily oxidized by heating in the air. Hydrochloric acid dissolves it easily, forming antimonous chloride and hydrogen sulphide.



On the other hand, a current of hydrogen sulphide passed into antimony solutions produces the antimonous sulphide as an *orange-red* precipitate, which by heating becomes like the natural form. These contrary reactions may be taken as an instance of the influence mentioned in § (c) page 23.

Antimonic Sulphide is an orange yellow body.

The chemical relation of antimony as a non-metal is shown more strongly in its sulphides than in its oxides. Both the sulphides act as sulphur anhydrides, and form with the metallic sulphides a series of sulphur salts.

KSbS_2 Potassium sulphantimonite,

is strictly comparable to

KNO_2 Potassium nitrite.

Antimonic sulphide, unlike the corresponding oxide, forms its salts upon the pattern of the orthophosphates.

Na_3SbS_4 Sodium sulphantimonate,

is analogous to

Na_3PO_4 Sodium orthophosphate.

Sodium sulphantimonate is used in photography under the name of *Schlippe's salt*.

BORON, B.

ATOMIC WEIGHT, 11. DENSITY, IN THE STATE OF GAS,

UNDETERMINED.

Sources. Occurs only in combination, either as boric acid in the discharges of steam in the volcanic region of Italy, or as sodium (or calcium) borate, found in certain lakes in Thibet, or upon the surface of the soil in western part of the U. S. Boron was first prepared by Davy in 1807.

Preparation. Boron is prepared by the action of potassium, sodium, or aluminium on boron trioxide.

Properties. Boron obtained by the action of sodium or potassium is an amorphous, olive green powder, insoluble in water, and combustible. When obtained by the action of aluminium, it is, if pure, in brilliant crystals closely resembling the diamond and next to it in hardness. A form of boron analogous to graphite has been supposed to exist, but it is probably in an impure condition.

Gen. Chem. Rel. Boron is a triad, and may be regarded as a link between the carbon and nitrogen group. Its compounds are not strictly analogous to those of any other element, but the existence of the amorphous and diamond forms, and its affinity for nitrogen, bring it more closely to carbon than to any other element.

Boric or Boracic Acid, H_3BO_3 . This exists, as already mentioned, in the steam jets discharged in some volcanic regions, and some of its salts occur as minerals. It can be prepared by dissolving borax (q. v.) in warm dilute sulphuric acid, and allowing the solution to cool. Boric acid forms pearly scales of a bitter

taste, soluble in water and alcohol, and very feebly acid. Heated to 248° F. (120° C.), it forms metaboric acid, HBO_2 , and on still further heating it is converted into boric anhydride, B_2O_3 , which fuses to a clear glass. When a solution of boric acid is boiled, some of the acid passes off in the steam. It forms salts called borates, many of which are irregular in composition. Boric acid is an antiseptic.

Tests. Boric acid has a feeble action on litmus, and turns turmeric paper to a brown-red color. The best test is the bright green color imparted to flame.

Exp. Shake a little boric acid with alcohol, pour the solution into a flat dish and ignite. The outer cone of the flame will be tinged with green.

Compounds of boron with nitrogen, chlorine, bromine and fluorine are known. Boron nitride is BN , but does not appear to be analogous to cyanogen CN .

SILICON, Si.

ATOMIC WEIGHT, 28.5. DENSITY, IN THE STATE OF GAS,

UNDETERMINED.

Sources. Never occurs in the free state, but is found very abundantly as oxide, SiO_2 , and as silicates. Discovered by Davy in 1807.

Preparation. By methods similar to those used for boron.

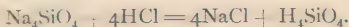
Properties. Silicon exists in three forms, amorphous, graphoidal and diamond, corresponding to those of carbon. When strongly heated in the air it burns, producing SiO_2 .

Gen. Chem. Rel. Silicon is a tetrad, and like boron, is related to carbon in many ways, especially in the capacity for assuming allotropic forms. Compounds have been obtained in which it has replaced carbon. Silicon also has some chemical relations to tin and titanium. It combines with many metals.

Silica, Silicic Anhydride, SiO_2 . This is a widely distributed body, occurring free as common sand, chalcedony, quartz, etc., and in combination forming silicates in great variety, of which clay, granite, feldspar and sandstones, are instances. A very large proportion of the solid substances in the earth's crust are forms or compounds of silica. Sand is simply a collection of small pebbles. Silica exist in the stems of grasses and in the teeth and bones of animals.

In its pure forms silicic anhydride is a colorless, nearly infusible, and insoluble solid, destitute of chemical activity. In nature, it often occurs in large six-sided crystals, either pure, or combined with minute quantities of metallic oxides, which give it distinct and often beautiful colors. In this way are formed agates, common amethyst, jasper, etc. It is converted into a silicate by fusion with sodium carbonate, and when lime, lead oxide, or other metallic oxides are mixed with the sodium silicates, we get the various forms of glass.

Silicic Acid, Orthosilicic Acid, H_4SiO_4 . This cannot be produced by direct union of water with the natural forms of silica, but is obtained when silicates are decomposed by acids. If the sodium silicate be treated with hydrochloric acid the following reaction occurs:



The silicic acid and salt remain in solution in the water in which the HCl was dissolved. By placing the solution in a vessel made of parchment paper and floating it in water, the salt passes out of the vessel and leaves a pure solution of silicic acid. This process is called dialysis. Solution of silicic acid is tasteless and feebly acid to litmus. By evaporation it forms a gelatinous mass which can be brought to the composition, H_2SiO_3 , (metasilicic acid), and by further heating gives the insoluble anhydride. The chemistry of silicic acid is complicated. It presents some of the characters of phosphoric acid in its power of forming different acids by taking up different proportions of water. Many of the natural silicates are decomposed and rendered soluble by the combined action of frost, and carbonic acid. In this way soils are formed and extensive geological changes ultimately produced.

Silicon combines with the halogens, forming bodies resembling the corresponding carbon compounds.

Silicon and fluorine have a strong affinity, and silicon fluoride, SiF_4 , is easily prepared by the action of hydrofluoric acid upon silicic acid, or any silicate. It is a colorless gas which is decomposed by water yielding gelatinous silicic acid and a double fluoride of silicon and hydrogen.



The latter body is often called hydrofluosilicic acid. Silicon combines with many metals, forming silicides, but many of these are of uncertain composition. Hydrogen silicide, H_4Si , the most interesting, is prepared by dissolving magnesium silicide in hydrochloric acid. It resembles hydrogen phosphide in taking fire spontaneously when impure. In its composition, it is analogous to marsh gas, CH_4 .

CLASSIFICATION OF THE ELEMENTS.

IT IS STATED on page 1, that the elements have generally been divided into two groups: metals and non-metals; but that the division is by no means satisfactory. It is to be regretted that the system has been so long retained by the leading chemical writers who alone have the authority to make changes in such matters, but it is evident from the disagreements among them, that modifications must soon be made. In the absence, therefore, of unanimity among the principal chemists, the minor writers and compilers have been left to their own inclinations, and the result is that scarcely any two works agree on the order in which the elements are described. The present compilation presents the peculiar feature of considering hydrogen as a metal, although on account of its frequent occurrence and important chemical relations, it has been described among the non-metals. In larger works, in which convenience and simplicity can be sacrificed to strictly scientific arrangement, hydrogen should be described with potassium and sodium; vanadium with nitrogen; tin and titanium with silicon; and oxygen with sulphur. It does not, however, appear possible to maintain a classification in two groups; many of the elements show a strictly intermediate condition. Under these circumstances it is allowable to attempt an arrangement in *three* groups—NON-METALS, METALLOIDS, METALS. These three groups may be divided into sub-groups, each containing a few elements whose chemical relations are essentially the same. In the following arrangement the sub-groups are separated by semicolons:

Non-metals,	O	S	Se	Te;	F	Cl	Br	I;	N	P;
	B	Si	C.							
Metalloids,	As	Sb	V;	Sn	Ti	Nb	Ta	W;	Os	
	Mo.									
Metals,	H	K	Na	Cs	Rb	Li;	Ba	Sr	Ca	
	Mg	Be	Zn	Cd;	Th	Y	E;	Al	Ga	
	In	Ce	La	Di;	Tl	Au;	Ag	Pb	Hg	
	Cu;	Ni	Co	Fe	Mn	Cr;	U;	Bi;		
	Ru	Pd	Rh	In	Pt;	Zr.				

Such arrangement can never show the true relations. The tabular form given on the opposite page, though complicated and necessarily imperfect, will show in a general way the relationships of the elements.

Since in the preceding descriptions, the characteristics of the groups have not been distinctly indicated, a brief account may be here given.

GENERAL CLASSIFICATION OF THE ELEMENTS.

Non-metals in **bold type** ; metalloids in *Italic* ; metals in Roman.

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1. *The Oxygen Group*, includes oxygen, sulphur, selenium and tellurium. They are electro-negative dyads and possess the power of forming two classes of compounds. United in low proportion, with the metals, they form bases; united with the non-metals or metalloids, and in high proportion with the metals, they form anhydrides. These bases and anhydrides can unite and form salts. Examples of these relations will be found on pages 34, 59.

2. *The Chlorine Group* includes chlorine, bromine, iodine and fluorine. They are electro-negative monads, and are distinguished as the only elements which can form salts without the aid of some member of the oxygen group. For this reason they have been called the HALOGENS, a word meaning salt-formers.

3. *The Nitrogen Group* includes nitrogen, phosphorus, arsenic, antimony and probably vanadium and bismuth. They are of uneven atomicity, generally triads or pentads, their electrical relations are intermediate in character, neither strongly positive, nor strongly negative. They form anhydrides, distinguished by the power of combining with water in two or more proportions, forming distinct acids, as noted in the descriptions of phosphoric antimonie and arsenic acid. The first four members of the group combine easily with hydrogen to produce bodies which have close analogies to the metals, as is explained under *ammonium* (*q. v.*)

4. *The Carbon Group* includes carbon and silicon. They are tetrads, and like the nitrogen group, their electrical relations are intermediate. Boron is usually classed here, but it is a triad and seems to be a connecting link between this and the nitrogen group. The existence of several allotropic modifications, amorphous and crystalline, of carbon, boron and silicon, is a point of much less importance than some writers have made it, as most of the elements present such variations.

The characteristics of each of the remaining groups will be given when the members of the group are described.

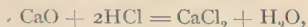
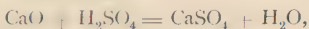
ACIDS, BASES AND SALTS.

THESE TERMS, introduced in the earlier days of chemical science, when the composition of bodies was but imperfectly understood are still retained, but with very vague and uncertain meaning. The old definition of an acid was a body having a sour taste, a power to affect vegetable colors, especially to turn blue colors red, and forming with the metals definite compounds called *salts*. Lavoisier had advanced the doctrine that all acids

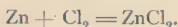
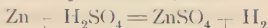
contained oxygen, but on the discovery of the sulphur acids and of the composition of hydrochloric acid, this view was abandoned. The study of the anhydrides, or as they were first called, anhydrous acids, showed that water was necessary to the acid condition, and thus was brought about the present view that hydrogen was the essential element of an acid. The explanation of these relations is given on pages 9, 10, 31, 33, 39.

Bases were defined to be compounds of the metals with oxygen, capable to uniting with an acid and neutralizing it. This definition was subsequently modified by including sulphur, selenium and tellurium as capable of converting metals into bases.

Salts were defined as bodies formed by the action of an acid upon a base. The reactions



are instances of such effects. The above definition, however, does not include the production of a salt by direct action of a halogen, or of an acid, upon a metal; thus:



Intimately connected with this subject is the meaning of the terms ACID, ALKALINE and NEUTRAL, as applied to the conditions of substances. If we add a drop of sulphuric acid to a solution of the coloring matter of purple cabbage, the color will change to red; by the addition of a small amount of soda, the color will be restored, and by further addition, will be changed to green. The soda is a base, it has combined with the acid and deprived it of its chemical activity. By this combination the soda has also been neutralized, and it is only by adding it *in excess*, that we can get its specific action on the color. Other coloring principles show similar effects.

Litmus is a red color that becomes blue on the addition of a base, and, of course, has the red color restored on the addition of an acid. It is usually sold in the blue condition, and is used either in solution in water or in the form of *litmus-paper*, strips of paper soaked in the solution and dried.

Cochineal is now much used as a substitute for litmus; acids turn it orange yellow, and bases turn it purplish.

It would seem that by the use of these tests, we could determine whether a substance was an acid, a base, or a salt, but unfortunately the reactions just given apply only to cases in which acids and bases of similar intensity are combined. When soda, which is a strong base, is united to sulphuric acid, which is a

strong acid, the compound is neutral, but the union of soda with a feeble acid like boric, produces a salt which is alkaline; and the union of copper oxide with sulphuric acid, gives a body which is acid to the tests mentioned above. Although these color reactions are of some importance in practical chemical operations, they have no value in determining the theoretical relations between acids, bases and salts.

Salts may be divided into four classes.

NORMAL SALTS, in which the positive ingredient consists of a single metal combined strictly according to its atomicities. Such are :

H_2SO_4	Hydrogen sulphate, (sulphuric acid).
Na_2CS_3	Sodium sulpho-carbonate.
KNO_3	Potassium nitrate.

MIXED SALTS in which two metals are present as positive ingredients: When one of the metals is hydrogen the body is usually called an *acid* salt.

HKCO_3	Acid potassium carbonate.
$\text{KNaC}_4\text{H}_4\text{O}_6$	Sodio-potassium tartrate.

It must be noticed that K_2CrO_4 is not a mixed salt, although it contains two metals; the chromium is not present as a positive metal.

DOUBLE SALTS in which two complete salts of either of the above classes unite to form a definite compound which is generally distinctly crystalline. Such are:

$\text{K}_2\text{SO}_4 + \text{Al}_2(\text{SO}_4)_3$	Potassium and aluminium sulphate.
$\text{BaCO}_3 + \text{CaCO}_3$	Barium and calcium carbonate.
$2\text{KCl} + \text{PtCl}_4$	Potassium and platinum chloride.

CONJUGATED SALTS in which a definite salt is united with a body not a salt. A great variety of these are known, and many are of uncertain composition. Two important classes may be recognized.

(a) *Oxy-salts*, called frequently *basic* or *sub-salts*, in which a base is united with the salt, thus:

$\text{Bi}(\text{NO}_3)_3 + \text{Bi}_2\text{O}_3$	Bismuth oxynitrate
$\text{SbCl}_3 + \text{Sb}_2\text{O}_3$	Antimony oxychloride.

(b) *Anhydro-salts*, called frequently *acid* or *bi-salts*, in which an anhydride is united with the salt, thus :

$\text{K}_2\text{CrO}_4\text{CrO}_3$	Potassium anhydrochromate.
$2\text{NaBO}_2 + \text{B}_2\text{O}_3$	Sodium anhydroborate.

Salts formed by one atom of a monad metal are called **MONO-BASIC**; formed by two atoms of a monad metal, **DIBASIC**; by three atoms **TRIBASIC**; by four **TETRABASIC**, thus ;

Metaphosphoric acid HPO_3 is monobasic.

Sulphuric acid H_2SO_4 is dibasic.

Sodium sulphantimonate Na_3SbS_3 is tribasic.

Potassium silicate K_4SiO_4 is tetrabasic.

General Characters of the Compounds of the Metals.

OXIDES are mostly insoluble in water and of little chemical activity. The oxides of the potassium group, called alkalis, are, however, very active substances. They are generally seen as hydrates and are very soluble in water.

SULPHIDES resemble the oxides, most of them being insoluble in water. Those of the potassium and barium groups are soluble.

CHLORIDES are normal salts, and are nearly all soluble in water; the most important exceptions to this rule being silver chloride, lead chloride and mercurous chloride.

BROMIDES and IODIDES agree with the chlorides.

CARBONATES are mostly insoluble in water, but are decomposed by most acids, and by heat. The metals of the potassium group form soluble carbonates, and those of the other metals are rendered soluble by excess of carbonic acid.

SULPHATES are soluble in water with the exception of those formed by barium, strontium and lead.

NITRATES are all soluble in water and decomposed by heat, generally leaving an oxide.

PHOSPHATES are mostly insoluble in water, but soluble in acids. The familiar forms are tribasic, and most of these are mixed salts. As instances we have:

Na_2HPO_4 Acid disodium phosphate.

NaH_2PO_4 Acid sodium “

$\text{Mg}.\text{Am}_2\text{H}_2(\text{PO}_4)_2$ Acid ammonio-magnesium phosphate.

Tests for Metals. Although it is a common custom to give, in connection with each metal, a list of tests, such tests are of no practical value unless performed in a systematic manner. Some metals have characteristic reactions by which they can be detected even in complicated mixtures, but the majority have to be searched for by a regular system. To explain this system in its full form does not come within the scope of this work, but the following is an outline, and presents the method used for examining compounds soluble in water, and supposed to contain one of the common metals.

To a portion of the solution in water, add HCl ; a precipitate shows either silver, lead or mercurous salts. If no precipitate is

formed, add to the same solution, H_2S ; a precipitate shows either lead, copper, bismuth, tin, gold, platinum, mercury, the metalloids arsenic and antimony, or possibly iron. If no precipitate forms, add to another part of the solution AmHO , AmCl and AmHS . A precipitate shows zinc, cobalt, nickel, iron, aluminium, chromium or manganese. If no precipitate forms, add to another part of the solution AmHO , AmCl , Am_2CO_3 . A precipitate shows barium, strontium or calcium. If no precipitate forms, add Na_2HPO_4 to the same liquid. A precipitate shows magnesium; no precipitate shows either potassium, sodium, or ammonium, for which special tests must be applied.

Flame Tests. Many metals give characteristic colors to flames. Such tests are very delicate, and when applied to pure substances, very satisfactory. When several metals are present, one color may conceal the other, and thus the test be incomplete. By passing the light through a prism the colors are separated and each may be recognized. The apparatus for this purpose is called a spectroscope. Observations with it show that many metals give out light which is made up of several colors.

Production of the Compounds of the Metals.

Although each metal requires special manipulation in the production of its compounds, yet it will be appropriate to give here an outline of the methods pursued in obtaining the more important forms.

METALS are obtained in the free state by reducing agents, or by the electrical current.

OXIDES are obtained (*a*) by direct oxidation, as in making lead oxide, by heating lead in the air; (*b*) by decomposing the carbonate, nitrate or hydrate by heat, as in making lime or copper oxide; (*c*) by direct action of nitric acid, as in making tin oxide.

HYDRATES are obtained either by the direct combination of the oxide with water, or more generally by the action of a caustic alkali upon a solution of a salt of the metal.

CARBONATES are usually obtained by the addition of a carbonate to a solution of the metal.

SULPHATES and **NITRATES** are usually obtained by the direct action of the acids.

Group 5. The Alkali Metals. This group includes K, Na, Li, Cs, Rb, and the quasi-metal Am, (NH₄). They are distinguished by high affinity for the non-metals, by forming very few insoluble compounds, and producing with oxygen, bases which are powerful chemical agents, and are known as the caustic alkalis. None of them exist free in nature. Hydrogen probably belongs here, but has been already described.

POTASSIUM, K. 39.

Sources and Preparation. Occurs in many rocks and soils; from these it is absorbed by land plants, in the ashes of which potassium carbonate is found. Large deposits of nitrate and chloride also occur. The metal is prepared by heating the carbonate with charcoal. Discovered by Davy in 1807.

Properties and Gen. Chem. Rel. A bluish-white soft metal, rusting quickly in the air and decomposing water rapidly. Specific gravity, 0.865. It is highly electro-positive and forms several oxides, only one of which is important.

Potassa, K₂O, is obtained by oxidizing potassium in dry air. It is generally seen as hydrate.

Potassium Hydrate, *Caustic Potassa*, KHO, is made by boiling potassium carbonate with slaked lime.



The solution is filtered from the insoluble CaCO₃, evaporated to dryness, the residue fused and cast in sticks. Caustic potassa is a white solid, very soluble in water, powerfully alkaline and corrosive.

Potassium Carbonate, K₂CO₃, *Salt of Tartar*. This is found in the ashes of land-plants being produced by the action of heat upon the compounds of potassium with organic acids. The ashes are treated with water, the solution thus obtained yields, on evaporation, the impure carbonate termed pearl-ash. Pure potassium carbonate is white, soluble in its own weight of water, alkaline and moderately corrosive.

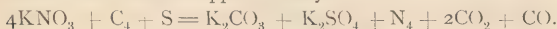
Acid Potassium Carbonate, KHCO₃, *Salaeratus*, is prepared by adding carbonic acid to the normal carbonate. It is a white, crystalline body, soluble in four times its weight of water. It is used in effervescing mixtures, but acid sodium carbonate has of late years substituted it to a great extent. It is often called *bicarbonate*.

Potassium Sulphate, K₂SO₄. This is a residue of some

manufacturing operations. It forms hard, colorless six-sided crystals, which are not very soluble in cold water.

Acid Potassium Sulphate, KHSO_4 , is also a by-product in certain operations. It is sour and strongly acid to test paper. It is often called bisulphate, and is used as a substitute and adulterant for cream of tartar.

Potassium Nitrate, KNO_3 , *Nitre*, *Saltpetre*, is found on the surface of the soil in India, and is prepared artificially by allowing nitrogenous matter to decay in the presence of wood ashes (containing K_2CO_3) and lime, and in a full supply of air. It appears that ammonia is first formed and then oxidized. The crude nitre, extracted by water, must be carefully purified, when it appears in large crystals, soluble in about four times its weight of cold and its own weight of boiling water. It melts below a red-heat and when further heated, decomposes, giving off oxygen and nitrogen, and leaving K_2O . Potassium nitrate is used in gunpowder and fireworks, as a source of oxygen. Gunpowder consists of about 75 parts nitre, 15 parts charcoal and 10 parts sulphur. The reaction is approximately :



The N , CO_2 and CO occupy at the moment of explosion about 1200 times the bulk of the powder, and the explosive action of gunpowder is due to this sudden expansion in volume.

Potassium Chlorate, KClO_3 . The method of manufacture is given on page 73. The salt crystallizes in flat tabular crystals. It melts at 806°F . (430°C .), at a little higher temperature gives off all its oxygen, leaving KCl . It is not very soluble in cold water. It is used largely as a source of oxygen, also in matches and fireworks, and as a medicine.

Exp. Some crystals of potassium chlorate, and a few pieces of phosphorus are put into a wine glass nearly filled with water, and sulphuric acid poured directly upon them by means of a funnel tube. The chlorate decomposes, furnishing various oxides of chlorine, which cause the phosphorus to burn brilliantly under the water.

Exp. Powdered sugar mixed with about three times its weight of potassium chlorate will burn when touched with a drop of sulphuric acid.

Exp. A mixture of two parts potassium chlorate with one part of potassium ferrocyanide and one part sugar makes *white* gunpowder, which explodes very easily and very violently. Only a small quantity should be made at once, and the ingredients should be powdered separately and mixed gently.

Potassium Chloride exists in sea-water and in the saline

deposits at Stassfurth, Germany. It resembles common salt. It forms an insoluble double salt with platinic chloride, $2\text{KCl} + \text{PtCl}_4$.

Potassium Bromide is made according to the method given on page 75. It forms cubical crystals soluble in water.

Potassium Iodide is prepared like the bromide, which it closely resembles, but is rather a finer white. It is easily soluble in water.

Flame Reaction. Potassium compounds give to flame a color which is a mixture of red and violet.

SODIUM, Na, 23.

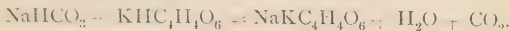
Sources. Common salt, NaCl , is the principal source. Sodium compounds are widely distributed, occurring even in common dust. The ashes of sea-plants contain sodium carbonate. Sodium was discovered by Davy in 1807.

Preparation and Properties. It is prepared in a manner similar to potassium, which it closely resembles, but is a little heavier and not so easily oxidized. Its chemical relations and the properties of its compounds are also much like those of potassium.

Sodium Hydrate, NaHO , Caustic Soda is prepared by a process similar to that used for caustic potash, using Na_2CO_3 instead of K_2CO_3 . Caustic soda is usually found in cylindrical sticks. It is soluble in water, is very strongly alkaline and corrosive.

Sodium Carbonate, Na_2CO_3 , Sal Soda, was formerly obtained from the ashes of sea-plants, but is now made by the action of chalk and charcoal upon sodium sulphate. It forms large crystals, having the composition $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$, and is often called *washing soda*. On exposure to air these crystals *effloresce*, that is lose water and fall to a white powder. It is very soluble in water.

Acid Sodium Carbonate, NaHCO_3 , Baking Soda. This body is produced in the same manner as acid potassium carbonate, which it closely resembles. It is now much used in effervescing mixtures like Seidlitz powders, and the common baking powders, which latter are usually a mixture of cream of tartar and baking soda. Alum and acid potassium sulphate are often used in the inferior grades as a substitute for the cream of tartar. The action of the powder is due to the sudden evolution of a large volume of carbon dioxide,



Acid sodium carbonate is a loose white powder soluble in water. It is often sold as *salaeratus*.

Sodium Sulphate, *Glauber's Salt*, Na_2SO_4 is a by-product in the manufacture of nitric and muriatic acids. See pages 43, 71. It forms large clear crystals which contain ten molecules of water of crystallization. They effloresce in dry air, and are remarkable for being more soluble in water at 93°F. (34°C.) than at any other temperature. The principal use of sodium sulphate is as a source of sodium carbonate.

Sodium Nitrate, NaNO_3 , is found in large beds in northern Chili, and termed Chili saltpetre. It is used as a manure, and also in the preparation of nitric acid. It cannot be used in gun-powder on account of its tendency to absorb water.

Sodium Chloride, *Common Salt*, NaCl , is too well known to need description. It occurs in thick beds in various parts of the world, and is also prepared from sea-water by evaporation or freezing; and from certain brine-springs by evaporation. It dissolves in about the same amount in hot and cold water.

Sodium Phosphates. The only important form is disodium acid phosphate Na_2HPO_4 , which is used in medicine and also as a test for magnesium.

Sodium Anhydroborate, $\text{Na}_2\text{B}_2\text{O}_4 + \text{B}_2\text{O}_3$, commonly called sodium baborate or borax, is found in certain lakes in Thibet, and in California. It is also made by melting sodium carbonate with boric acid. It forms hard crystals which dissolve in about twelve times their weight of water, and form an alkaline solution. Borax is much used as a solvent for metallic oxides especially in blow-pipe analysis. It is used for cleaning metals in soldering.

Sodium Silicate, made by fusing sand or pulverized quartz with an excess of sodium carbonate, constitutes *soluble glass*, which dissolves in boiling water. It is used as a cement, and in soaps.

Sodium Thiosulphate, $\text{Na}_2\text{S}_2\text{O}_3$, commonly called *hypo-sulphite*, is prepared as explained on page 67. It forms large crystals very soluble in water. The solution possesses the power of dissolving most of the salts of silver, except Ag_2S , which are insoluble in water. It is used in photography.

Sodium Sulphite, Na_2SO_3 , is used as a substitute for sulphurous acid for preventing fermentation.

Flame Reaction. Sodium gives to flame a strong yellow color.

LITHIUM, Li , 7. Discovered by Arfvedson in 1817. It is found in many substances but only in small quantity. Its principal sources are some rather rare minerals. The metal resembles potassium and is next to hydrogen, the lightest metal known.

Specific gravity, 0.593. Its salts resemble those of potassium and sodium, but lithium carbonate is but sparingly soluble in water. Lithium compounds have been used in gout and similar chronic diseases. Many spring waters contain traces of lithia. The medicinal action of such waters has been supposed to be due to the lithia, and much nonsense and quackery have been developed in connection with the analysis of them.

Flame Reaction. Lithium imparts a crimson color to flame.

CÆSIUM, Cs, 133, and RUBIDIUM, Rb 85.4 were discovered by Bunsen and Kirchhoff, in 1860, by the spectroscope. They exist only in small quantity, in some mineral waters and in a few plants. They are strongly positive and closely resemble potassium. Cæsium gives a blue color to flame; rubidium a dark red color.

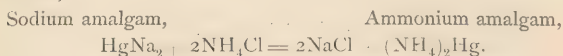
AMMONIUM.

In the description of ammonia gas, NH_3 , page 42, mention is made of its affinity for acids, and of the strongly alkaline properties of its solution in water. Long before the composition of the solution was understood, it had received the name of *volatile alkali*, to indicate its resemblance to potassa and soda, which were called fixed alkalies. Berzelius was the first to suggest that the compounds produced by ammonia were formed in the same manner as those of potassium and sodium, the metal being represented by a compound body, NH_4 . In this way $\text{NH}_3 + \text{HCl}$ would produce NH_4Cl ; $\text{NH}_3 + \text{H}_2\text{O}$, would produce NH_4HO . NH_4 is called ammonium; its atomicity is of course, monad; like monad metals, it combines with one atom of chlorine and replaces the hydrogen of acids. In short, this body is capable of acting exactly like potassium or sodium, and, while in strictness it cannot be considered a metal, it has received a name which terminates in the syllable *um*, characteristic of the metals. The following formulæ show the comparison between the salts of potassium and ammonium:

KCl	Potassium chloride.
NH_4Cl	Ammonium chloride.
K_2SO_4	Potassium sulphate.
$(\text{NH}_4)_2\text{SO}_4$	Ammonium sulphate.
KNO_3	Potassium nitrate.
NH_4NO_3	Ammonium nitrate.
KHO	Potassium hydrate.
NH_4HO	Ammonium hydrate.

As the expression NH_4 makes some confusion in formulæ, it is convenient to use the symbol Am ; ammonium chloride will thus be AmCl ; ammonium sulphate, Am_2SO_4 . The ammonium theory received considerable support when proposed, from the following experiment:

If a piece of sodium be put on the surface of some mercury, and a drop of water added cautiously, an alloy called sodium amalgam will be formed. If this be dropped into a strong solution of ammonium chloride, a remarkable increase of bulk of the amalgam will take place. It was, at first, supposed that a compound had been formed between the mercury and NH_4 by the following reaction,



The property of forming an amalgam with mercury was regarded as proof of the metallic nature of NH_4 . Subsequent research has shown that the supposed ammonium amalgam is nothing but mercury inflated with gas bubbles. The mass rapidly diminishes in bulk, giving off a mixture of NH_3 and H .

Although this experiment is no longer of any weight, the general chemical reasons are sufficient to maintain the ammonium theory and it is now usual to describe the ammonium compounds with those of the alkali metals.

A compound body which has a power of acting like a element is called a quasi-element (Lat. *quasi*, as if), or radicle. Ammonium is a quasi-metal; cyanogen, CN , is a quasi-non-metal.

The termination "um" is changed into "a" when the metal is combined with oxygen. Accordingly, ammonium will form ammonia $(\text{NH}_4)_2\text{O}$ or Am_2O , corresponding to potassa K_2O or soda Na_2O . A little confusion occurs here, in consequence of the name ammonia applying also to NH_3 , which for distinction, should be called ammonia gas, or better *amine*. In ordinary chemical language, "ammonia" always refers to the oxide Am_2O , or the hydrate.

Ammonium Compounds. The sources of ammonia gas, the properties of its solution in water and its tests have been given on pages 41 and 42. It remains to describe some of the salts.

Ammonium Sulphide, Am_2S , is produced by saturating hydrogen sulphide with ammonia solution. By half saturation we get AmHS , ammonium sulphhydrate. Both compounds are yellowish liquids of disgusting odor. They are much used in analytical chemistry.

Ammonium Carbonate, Am_2CO_3 , is produced by the decomposition of organic matter, especially in sewer water. It is

made on the large scale by heating ammonium chloride with chalk. The theoretical reaction would be



The Am_2CO_3 , however, decomposes into Am_2O , which escapes and leaves $2\text{Am}_2\text{CO}_3 \rightarrow \text{CO}_2$, ammonium anhydro-carbonate, (see page 96.) often called sesquicarbonate, or smelling salt. It is a white body, soluble in water, and smelling strongly of ammonia. By exposure to air it is converted into acid carbonate, AmHCO_3 .

Ammonium Nitrate, AmNO_3 , is made by saturating nitric acid with ammonia. It is a white solid, very soluble in water. Its chief use is for making nitrous oxide, (q. v.)

Ammonium Sulphate, Am_2SO_4 , is obtained by boiling gas liquor and passing the vapors into sulphuric acid. It is used as a fertilizer, and in the manufacture of alum.

Ammonium Phosphates. Of these, the most important is microcosmic salt, AmNaHPO_4 . It is used in blowpipe analyses. Various ammonium phosphates exist in manures, and are valuable fertilizers.

Ammonium Chloride, *Sal Ammoniac*, AmCl , is prepared as described on page 41. It is a white solid, crystallizing in cubes, and is very soluble in water. It volatilizes without fusing, and is apparently decomposed into HCl and NH_3 . It has many uses in the arts, in medicine and in analytical chemistry.

Ammonium Bromide and *Ammonium Iodide* are used in photography and medicine.

Flame Reaction. Ammonium salts give a greenish color to flame. In the spectroscope the colors are seen to be those of the components, nitrogen and hydrogen.

Group 6. Alkaline-earth Metals. These form oxides, which are slightly soluble in water but much less caustic, or corrosive than the alkalis proper. Their sulphates, carbonates and phosphates are practically insoluble. None of the metals occur in the free state in nature.

BARIUM, Ba, 137.

Sources, etc. Occurs principally as sulphate and carbonate. Discovered by Davy in 1808. It is prepared from the chloride by electrical decomposition, and is a pale yellow, easily oxidized metal, of a specific gravity of 4.

Barium Oxide, Baryta, BaO , is obtained by heating the nitrate. It easily takes water, forming barium hydrate, BaH_2O_2 , which is soluble in about 20 parts of water, the water being rendered turbid by even very small amounts of carbonic acid, and hence used^r as a test.

Barium Dioxide, BaO_2 , is formed by heating BaO in a current of air. It is used in making hydrogen dioxide.

Barium Carbonate, BaCO_3 , is found in nature as *witherite*, and is also made artificially. It is insoluble in pure water, and used in some analytical operations.

Barium Sulphate, BaSO_4 *Barytes, Heavy Spar*, is found abundantly as a mineral, often forming the *gangue*, or rock surrounding metallic veins. It is a very heavy, white, insoluble solid, often finely crystallized. It is used as a substitute and adulterant for white lead, and also to adulterate various other articles. When barium salts are mixed with a sulphate, the barium sulphate is thrown down. Heated with carbon, it becomes a soluble sulphide and can in this way be used as a source of the other barium salts.

Barium Nitrate, $\text{Ba}(\text{NO}_3)_2$ is used as a test solution for sulphates, and in making green fire.

Barium Chloride, BaCl_2 , is also used as a test.

Flame Reaction. Barium communicates to flame, a yellowish green color, which the prism shows to contain several distinct shades.

STRONTIUM, Sr , 87.5, resembles barium closely in its compounds and chemical relations. It occurs as sulphate *celestine* and carbonate *strontianite*. The metal is prepared similarly to barium. It was discovered by Davy in 1808.

Strontium Sulphate, SrSO_4 , is very slightly soluble in water.

Strontium Nitrate, is used in making red fire.

Flame Reaction. Strontium compounds give to flame a crimson tint, which is a mixture of red, orange and blue.

CALCIUM, Ca , 40.

Sources. Occurs abundantly as sulphate carbonate, phosphate and fluoride. Discovered by Davy in 1808.

Preparation and Properties. Calcium is prepared by decomposing the chloride either by electricity or by sodium. It is a light yellow, hard metal, not brittle, oxidizing easily.

Calcium Oxide, *Quicklime*, CaO , is obtained by heating the carbonate to redness.



Quicklime, (quick means alive or active), is a white, infusible solid, which neutralizes the most powerful acids, and combines with water with great energy, forming

Calcium Hydrate, *Slaked Lime*, CaH_2O_2 , a soft white caustic powder, slightly soluble in cold water, (about 9 grains to the pint) less so in hot. The solution known as *lime water*, is used as a test for free carbonic acid, which produces a precipitate in it.

Slaked lime mixed with sand constitutes mortar. The cause of the hardening of mortar is not definitely known; calcium carbonate and silicate are formed, but only in small amounts.

Lime is used in agriculture to assist in decomposing the silicates in the soils.

Calcium Carbonate, CaCO_3 , occurs abundantly and in a variety of forms. It is the chief constituent of shells and of coral. In a non-crystalline condition it is seen as *chalk*, *marble* and *limestone*; in crystals it forms *Iceland spar* and *Arragonite*. It can be prepared artificially by adding sodium carbonate to calcium chloride,



Calcium carbonate is a white solid, insoluble in water and decomposed by a red-heat. It dissolves in water containing carbon dioxide, for which reason most spring and river waters contain it. When present in an amount more than a grain or two to the gallon, a *hard* water is formed, which has the property of curdling soap and preventing the formation of a lather. This is due to precipitation of insoluble calcium salts, formed from the fat-acids of the soap. Boiling the water will expel the carbon dioxide and precipitate the calcium carbonate, thus softening the water. The same result is attained by adding clear lime water, which combines with the carbon dioxide. Water which can be softened by these methods, is said to be temporarily hard. It is probable that the calcium carbonate exists in the water as an anhydro-carbonate, analogous to borax; (see page 96.) The excess of carbon dioxide may also be expelled by exposure of the water to air, and the calcium carbonate will then be deposited. Such an action occurs for instance, in caves, forming *stalactites* and *stalagmites*.

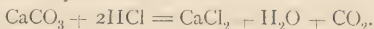
Calcium Sulphate, CaSO_4 , usually occurs as a mineral, crystallized with $2\text{H}_2\text{O}$, constituting selenite, gypsum and alabaster, sometimes, however, anhydrous. It is soluble in about 400 times its weight of cold water. It is a frequent ingredient of spring and river water, causing the same effect of hardness mentioned above, but as the sulphate does not owe its solubility to carbon

dioxide, boiling, except for a long time, does not soften the water, and hence the condition is called *permanent* hardness. When the crystallized mineral is heated moderately, it loses its water of crystallization and becomes a soft white powder, (plaster of Paris) which when mixed again with water, re-absorbs it, and becomes a hard mass, expanding slightly in bulk, and thus very suited for taking casts of any object.

Calcium Phosphate, Ca_3PO_4 occurs in bone and in modified form in some mineral deposits. Its chief use is in fertilizers and in the manufacture of phosphorus and its compounds. It is insoluble in water.

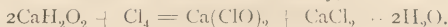
Calcium Hypophosphite, $\text{Ca}(\text{PH}_2\text{O}_2)_2$, is prepared by boiling phosphorus with lime, as indicated on page 80. It is used in medicine. By substituting barium hydrate for the lime, barium hypophosphite is formed, from which *hypophosphorous acid* can be obtained as a syrupy liquid, easily decomposed, and having powerful reducing action.

Calcium Chloride, CaCl_2 , is easily obtained by dissolving the carbonate in hydrochloric acid.

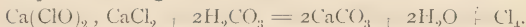


It is very soluble in water, and the solution, on evaporation forms crystals of the composition $\text{CaCl}_2 \cdot 6\text{H}_2\text{O}$, which melt and become anhydrous at about 450°F . (230°C .) The anhydrous salt has a powerful affinity for water, and is used for drying gases. It also absorbs ammonia gas and combines with alcohol.

Calcium Hypochlorite, *Bleaching Powder*. This body is produced by passing chlorine into slaked lime, keeping the mixture cool. The reaction should be in theory,



The exact composition of the commercial bleaching powder is, however, still undetermined; it appears to contain some unchanged calcium hydrate, but is probably mainly a combination of calcium hypochlorite, $\text{Ca}(\text{ClO})_2$, with calcium chloride, CaCl_2 . Bleaching powder, when in good condition, is a loose, dry, white powder, with a faint and not disagreeable odor. If it smells of chlorine, it is in bad condition. It dissolves in water. The solution possesses strong bleaching and deoxidizing powers, for which purposes it is largely used. Acids, even carbonic acid, decompose it, setting chlorine free.



The commercial salt is often erroneously called *chloride of lime*.

Calcium Fluoride, CaF_2 , *Fluor Spar*, is the principle source of the fluorine compounds. It is found as a mineral, often in

very fine colored crystals which become luminous when heated.

Flame Reaction. Calcium compounds give to flame a reddish color which is a mixture of orange, green and a faint blue.

Group 7. Earth Metals. These are mostly quite rare, and never occur in the free state. Most of them form but one oxide, in which the metal acts as a double tetrad, as described on page 16. Cerium, and possibly didymium form two oxides. All the oxides are insoluble in water and destitute of caustic action. The group is also noted for its tendency to form double salts with the salts of the alkali metals.

ALUMINIUM, Al, 27.5.

Sources. Very abundant as silicate, constituting clay and many common rocks. Most building materials are mixtures of aluminium silicate with other silicates. Discovered by Woehler, in 1828.

Preparation. By decomposing the chloride with sodium, and by several other methods.

Properties. Aluminium is a white, not very lustrous metal, easily beaten into plates or drawn into wire, sonorous and very light; specific gravity, 2.6. It tarnishes slightly in the air, and dissolves rapidly in hydrochloric acid and in caustic alkalies, but not in nitric acid. It melts at 842°F. (450°C.) Its alloys are very valuable; that with copper, aluminium bronze, has the color of gold, and keeps well in the air.

Aluminium forms but one series of compounds, which possess strong analogies to the tetrad series of iron salts, hence, aluminium is regarded also as a tetrad. Like iron, it forms compounds by the joint action of two tetrad atoms, which act as a hexad, as explained on pages 17 and 19.

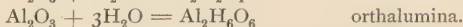
Aluminium Oxide, Al_2O_3 , occurs naturally as *corundum* which, when crushed, constitutes *emery*; finely crystallized, as the *ruby* and *sapphire*. It can be prepared by heating ammonia alum; it then forms a white powder. In all the anhydrous forms it is absolutely insoluble in water, and almost so in acids and alkalies.

Aluminium Hydrate, $\text{Al}_2\text{H}_6\text{O}_6$, may be regarded as formed by the reaction $\text{Al}_2\text{O}_3 + 3\text{H}_2\text{O} = \text{Al}_2\text{H}_6\text{O}_6$, but it has not yet been made in this manner. It is usually obtained by adding ammonia to a solution of the sulphate.



This hydrate is a gelatinous, white mass, easily soluble in acids and alkalies, and has a strong affinity for organic matter; with organic colors it forms precipitates called lakes. In the art of dyeing, alumina is used as a *mordant*, that is, a material to cause the color to adhere firmly to the fabric.

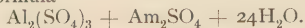
Aluminium oxide appears to possess a property somewhat similar to phosphoric acid, P_2O_5 , of forming different hydrates by taking up different proportions of water.



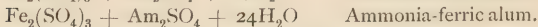
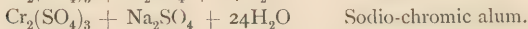
These are found as minerals, the first known as diaspore, the second as gibbsite.

Aluminium Sulphate, $\text{Al}_2(\text{SO}_4)_3$, is now largely manufactured for use in dyeing and in other operations. It is prepared either by the action of sulphuric acid upon clay, nearly free from iron, or by the roasting of a mineral termed *alum shale*, which contains iron disulphide, FeS_2 , and aluminium silicate. The roasting converts the sulphur into sulphuric acid which then forms aluminium sulphate and iron sulphate. These are dissolved in water and separated by crystallization. $\text{Al}_2(\text{SO}_4)_3$ usually forms a white crystalline mass, having an acid reaction.

Alum. The word alum, like the term alcohol, long applied to a single substance, has now come to refer to a group of substances characterized by similarity of composition and chemical relations. The alums are a series of double sulphates. Common alum has the formula



The aluminium in this compound may be replaced by chromium, iron, manganese, and practically by all the metals which have the power to act as double tetrads; (see page 16.) The potassium may be replaced by any metal of its group, and by silver. In this manner we get a series of salts of which only a few instances need be given:

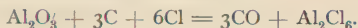


The particular point about these compounds is that they all contain the same amount of water of crystallization, and all crystallize in octahedra. Bodies that are similar, both in composition and crystalline form, are called **ISOMORPHOUS**.

Common Alum, Ammonia Alum, crystallizes in large, clear octahedra, which dissolve easily in water, the solution being acid to test paper, and strongly astringent.

When alum crystals are gently heated, they swell up, lose their water of crystallization, and fall to a soft white powder, burnt alum.

Aluminium Chloride, Al_2Cl_6 , is prepared by heating alumina and charcoal in a current of chlorine,



Glass, Pottery and Porcelain. These important articles are mixtures of various silicates, especially aluminium, calcium and sodium silicates.

Window and plate-glass contain sodium and calcium silicates. Bohemian glass contains potassium and calcium silicates. Flint glass or crystal contains potassium and lead silicates. Common green bottle glass contains sodium, calcium, iron and aluminium silicates.

Oxide of lead increases the lustre of the glass, as well as its weight and fusibility.

Colored glasses are produced by the use of various oxides. Ferrous oxide, FeO , produces a deep green (bottle glass), and manganese dioxide communicates a purple tint. It is difficult to obtain materials entirely free from iron, which gives a green tinge; for a colorless glass, manganese dioxide is added.

Pottery and earthenware are made of clay, moulded while wet and then baked at a temperature which renders it no longer plastic. The glaze is a fusible sodium silicate made from a mixture of sand and salt. Lead silicate is also used. Porcelain is a mixture of feldspar (aluminium and potassium silicate), sand and *kaolin*, a hydrated aluminium silicate. These when mixed with water, moulded to shape and heated to a high temperature, form a porous ware called *biscuit*. This is glazed by fusing on the surface various silicates.

The colors on pottery and porcelain are mostly mixtures of sand and metallic oxides.

Flame Reaction. Aluminium compounds give no color to common flames. When intensely heated, by electric sparks, a number of tints, chiefly red and blue, are produced.

The remaining members of the aluminium group need very short description.

CERIUM, Ce. LANTHANUM, La and DIDYMIUM, Di, occur together in a few rare minerals. They form oxides which are analogous to alumina, and highly insoluble oxalates. Solutions of didymium salts when examined by the spectroscope, show several *absorption bands*, that is, points at which the light is interrupted.

INDIUM. It exists in certain ores of zinc obtained at Freiberg, Saxony. It was discovered in 1863 by Reich and Richter. It gives when heated a violet blue flame, which is a mixture of indigo and blue.

GALLIUM, Ga, 68.9, exists in small quantities in various zinc ores. It has only been obtained in small quantity. Discovered by Lecoq de Boisbaudran in 1875. Gallium melts at 86° F. (30° C.) and will therefore melt by the heat of the hand.

Group 8. Zinc Group. This includes zinc, magnesium, cadmium and glucinum. They are never found free but are tolerably easily reduced from their compounds. They each form but one definite oxide which is insoluble in water, not caustic, but capable of forming well marked salts.

ZINC, Zn, 65.

Sources. Exists rather abundantly as sulphide (blende), carbonate, (calamine), silicate, (electric calamine), and as oxide. Zinc was known in the 13th century.

Preparation. The ores are first converted into oxide by roasting in the air, and the oxide is then heated with charcoal in an earthen retort, or crucible so arranged that the metal may distil into a receiver.

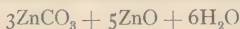
Properties. Zinc is a hard bluish-white metal, generally decidedly crystalline, Sp. Gr. 7.14. It melts at 770° F. (410° C) and distills at about a red-heat. It is brittle at ordinary temperatures, but becomes soft and easily beaten out at between 212° and 302° F. (100° and 150° C.), and at a higher temperature again becomes brittle. Zinc oxidizes slightly in moist air, and when highly heated, burns with a greenish white flame, producing ZnO. Acids and strong alkalies dissolve it. The uses of the metal are numerous. It is employed in making several important alloys, as brass and gun-metal, which contain copper and zinc; and German silver, which contains copper, zinc and nickel. Galvanized iron is simply iron covered with a layer of zinc by dipping it in a bath of the melted metal. Commercial zinc is very likely to contain arsenic.

Zinc Oxide, ZnO, *Zinc White*, is easily made by burning zinc. It is a soft powder, yellow when hot, white when cold. It is used as a paint, as application in surgical dressings, and as a face powder. It dissolves easily in acids.

Zinc Hydrate, ZnH_2O_2 , is produced by adding ammonia to a solution of a zinc salt. It is a white precipitate soluble in acids and alkalies.

Zinc Sulphide, ZnS , occurs as a crystalline mineral called *blende*. It is produced artificially as a white gelatinous mass as described on page 60.

Zinc Carbonate, ZnCO_3 , *Calamine*, exists as a mineral. When zinc sulphate and sodium carbonate are mixed, an oxy-salt,



is thrown down.

Zinc Sulphate, ZnSO_4 , *White Vitriol*, is made either by dissolving the metal in sulphuric acid or by oxidizing the sulphide. It forms white crystals, having the formula $\text{ZnSO}_4 + 7\text{H}_2\text{O}$. They are soluble in water, have an acid reaction and act as an emetic.

Zinc Chloride, ZnCl_2 , is made by dissolving zinc scraps in hydrochloric acid. It forms white masses, which absorb water rapidly from the air, (*deliquesce*) and make a strong solution. Zinc chloride is a powerful corrosive, coagulates albuminous matter, and is used as a preservative in anatomical preparations, also as an application to exposed tooth pulps, etc. When a strong solution of zinc chloride is mixed with zinc oxide, the two combine and form a hard white insoluble mass, which is used as a filling for teeth.

Zinc Phosphate, or rather *Oxy-phosphate*, made by mixing zinc oxide with phosphoric acid, and consisting of zinc phosphate united with zinc oxide, has come into use lately as a substitute for the oxychloride in filling teeth.

Flame Reaction. Zinc burns with a bluish flame, which is mixture of red and blue.

MAGNESIUM, Mg, 24.

Sources. Occurs as carbonate (dolomite), silicate (talc and soapstone) also as hydrate and chloride. Most natural water contains magnesium compounds. The metal was discovered by Davy in 1808.

Preparation and Properties. It is obtained by heating the chloride with sodium. It is a bright metal easily beaten into sheets or drawn into wire. Sp. Gr. 1.74. When strongly heated in the air it burns with a very bright light, producing MgO . Magnesium compounds often cause *hardness* in water, similar to that produced by calcium salts.

Magnesium Oxide, MgO , *Magnesia* is usually obtained by heating the carbonate to redness. It is a light white powder very feebly soluble in water and neutralizing acids. Magnesium hydrate is also known.

Magnesium Carbonate, MgCO_3 , occurs as a mineral. It is usually obtained by adding sodium carbonate to magnesium sulphate. In this case an oxycarbonate is formed, which is known as *magnesia alba*.

Magnesium Sulphate, MgSO_4 , *Epsom Salt*. This body is obtained on the large scale by the action of sulphuric acid upon natural magnesium carbonate and silicate. It forms colorless crystals *isomorphous* (see page 110.) with zinc sulphate. As thus crystallized, the salt has the composition $\text{MgSO}_4 + 7\text{H}_2\text{O}$. It is very soluble in water.

Magnesium Chloride, MgCl_2 , resembles zinc chloride in some respects, but does not have the same corrosive action.

Flame Reaction. Magnesium gives no color to a gas flame; with the electric spark it gives a green light.

CADMIUM, Cd , 112. Occurs in various zinc ores. It is extracted in the same manner as zinc, and forms a silver white crystalline metal, sp. gr. 8.6, melting at 500°F . (260°C .) and is nearly as volatile as mercury. It is easily dissolved by ordinary acids. Its salts are of little interest, except the sulphide. CdS is obtained as orange yellow precipitate by passing hydrogen sulphide into a solution of cadmium.

Beryllium or Glucinum, Be , 9.3, occurs in beryl, emerald and in a few other minerals. It is rare and unimportant. The metal is white, and malleable. Specific gravity, 2.1.

Thorium, Th , 231.5, was discovered by Berzelius in 1828, in a silicate called *thorite*.

Yttrium, Y , 93, was discovered by Gadolin in 1794, and *Erbium*, E , 169, was discovered by Mosander in 1843. They exist in a mineral called *gadolinite*, found at Ytterby, Sweden.

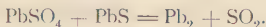
These three metals have been so little studied on account of their rarity that it is as yet not possible to classify them definitely.

Group 9. The Lead Group. This includes lead, copper and mercury. They form a basic oxide practically insoluble in water, and an insoluble chloride. By this latter property, silver and thallium are connected with the group. Copper, silver and mercury occur in the free state in nature. All the five metals resist the action of the air.

LEAD, Pb, 207.

Sources. Occurs abundantly as sulphide, (*galena*) and in small quantity as carbonate, sulphate and phosphate. It was known to the ancients.

Preparation. The sulphide is the common ore. It is roasted in a free supply of air, by which a portion is converted into sulphate. The mixture is then highly heated and the following reaction occurs :



Properties. Lead is a soft bluish-white metal, not capable of being made into very thin sheets, or wire. It resists the action of the air, and even of some strong acids, for which reason it is often used in chemical apparatus when glass cannot be used ; as in the manufacture of sulphuric acid. Pure water free from air, has no action on lead, but aerated water soon oxidizes and dissolves it in small quantity. The presence of sulphates and phosphates interferes with this action, because they precipitate insoluble lead compounds. The composition of a water is thus a matter of some importance, when the supply comes through lead-pipes ; and the presence of a few grains of calcium or sodium sulphate to the gallon, is a good preventative of lead poisoning. Lead melts at 617° F. (325° C.) and boils at a white heat. Specific gravity 11.5.

It forms some important alloys. *Type metal* contains 4 parts lead, and 1 part antimony ; solder about equal parts of lead and tin ; pewter 1 part lead, and 4 of tin.

Lead Monoxide, Litharge, Massicot, PbO , is usually made by heating lead in the air. It is a yellowish or reddish powder, slightly soluble in water and neutralizing the most powerful acids. It fuses at a red heat ; and in this condition combines easily with silica for which reason it is often used in glazing earthenware, but such glaze is easily attacked by acids, and may give rise to lead poisoning. Lead oxide is used in paints and cements.

Lead Dioxide, Puce or Brown Oxide, PbO_2 , may be obtained by the action of nitric acid upon red lead. It is a brown powder insoluble in water, and having some of the characters of an anhydride. It is a powerful oxidizing agent,

Red Lead, Minium, usually Pb_3O_4 , is a mixture of PbO and PbO_2 . It is obtained by heating litharge in the air for some hours. It forms a bright red powder not constant in composition. It is used as a coloring material, and in the manufacture of glass.

Lead Sulphide, PbS , is abundant as a mineral, forming large cubical lead-colored crystals. Lead sulphide is easily made artificially by passing hydrogen sulphide into lead solutions, thus :



Lead Carbonate, PbCO_3 , *White Lead*, occurs as a mineral, but is now made artificially on a very large scale, for use in paints. The most used method is known as the Dutch process. Thin sheets of lead are loosely rolled up and set in earthen jars, at the bottom of which some vinegar is placed. The jars are then piled in rows and covered with spent tan, but the air is not entirely excluded. The tan decomposes producing heat and evolving carbon dioxide. A lead oxyacetate is first formed and then converted into carbonate. The process requires some weeks, and the product is then ground, and for use is mixed with fixed oils. Lead carbonate is a white, opaque powder, insoluble in pure water. The white lead of commerce is an oxy-salt of varying composition, approximately $2\text{PbCO}_3 + \text{PbO} + \text{H}_2\text{O}$.

Lead Sulphate, PbSO_4 , is thrown down when lead salts are mixed with any sulphate. It is a white insoluble powder.

Lead Chloride, PbCl_2 , is not very soluble in cold water, so that when a chloride is added to a lead salt, a precipitate often occurs. It forms slender crystals, which are tolerably soluble in boiling water.

Lead Iodide, PbI_2 , is easily produced by adding potassium iodide to lead solution. It forms a bright yellow powder, sparingly soluble in cold water. If dissolved in boiling water, the solution on cooling will deposit the iodide in crystals.

Flame Reaction. Lead compounds give a pale green color to gas flame; with the electric spark a mixture of violet, green and yellow is produced.

COPPER, Cu, 63.

Sources. Occurs native in large masses, also abundant as sulphide, copper pyrites; and as oxide, silicate and carbonate. The metal was known to the ancients.

Preparation. The native copper, of course, requires very little preparation. The sulphide and carbonate are decomposed by roasting and successive meltings.

Properties. Copper is distinguished by its red color. It is heavy, specific gravity, 8.9, hard and can be worked into thin plates or wire, melts at 1996° F. (1091° C.) It conducts heat and electricity very well, and resists the action of the air, but is slightly oxidized and dissolved by acids when in contact with air. Even sea-water and the acids of fruits will produce this effect, and hence the danger of using copper vessels for kitchen purposes. The metal furnishes us with some valuable alloys; brass, gun-metal, etc. An alloy of copper with hydrogen Cu_2H_2 , has attracted some attention on account of its supposed theoretical relations. Copper compounds form two sets of salts; in the most important, the *cupric* series, the metal is dyad; in the other, *cuprous*, two atoms of copper are supposed to act together as a double atom. The condition is usually explained by graphic formulæ, (see page 18), thus:

Copper in cupric salts,

—Cu—

Copper in cuprous salts,

—Cu—Cu—

The cuprous salts are mostly colorless; the cupric salts are green, or blue.

Copper Monoxide, Black Oxide, CuO , is prepared by heating copper in air, or by roasting the nitrate. It is a heavy black powder dissolving in acids. Heated in a current of hydrogen, it is easily reduced to metallic copper. Copper oxide is used in organic analysis. Copper hydrate, CuH_2O_2 , formed when copper salts are mixed with an alkali, is a bluish-green mass, dissolving in ammonia, producing a clear deep blue liquid of complicated composition. With potassa and soda no solution occurs except in the presence of certain organic bodies, especially sugar, when a clear blue solution is also formed. If such solution is boiled, the cupric hydrate is changed to cuprous, which is precipitated as a red or orange powder. This reaction is a useful test for sugar (*q. v.*)

Copper Sulphide, CuS , *Copper Pyrites*, occurs as a mineral, and is easily produced as a black precipitate by mixing copper solution with hydrogen sulphide.



Copper Carbonate, CuCO_3 , appears not to be known in the pure state. Various oxy-carbonates, malachite and azurite, exist as fine minerals, and similar compounds are obtained by the addition of carbonates to copper salts. The natural forms are used for ornamental articles, the artificial form for paints.

Copper Sulphate, CuSO_4 , *Blue Vitriol*, *Blue Stone*, is formed by dissolving the metal or its oxide in sulphuric acid. It forms

large blue crystals, soluble in water, and having the composition $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$. When heated, the water of crystallization is driven out, and the salt becomes a soft white powder.

Copper Nitrate, $\text{Cu}(\text{NO}_3)_2$, is a residue from the preparation of nitric oxide. It has very little importance.

Copper Chloride, CuCl_2 , is in green crystals, soluble in water. It forms a number of double salts. An oxychloride is used as a paint.

Copper Arsenite, Scheele's, or Paris green, Cu_2HAsO_3 , is a bright green powder, obtained by mixing an alkaline arsenite with cupric sulphate. It is used for killing potato bugs, and also as a color. It is a violent poison. A compound of acetate and arsenite is known as Schweinfurth green.

Cuprous Salts. These are of little importance. *Cuprous Oxide*, Cu_2O , has already been mentioned as the result of the action of sugar on a mixture of caustic alkali and cupric hydrate; *cuprous chloride*, Cu_2Cl_2 , is a white solid, insoluble in water. The cuprous salts are mostly easily converted into cupric.

Flame Reaction. Copper gives a green tint to flame; with the electric spark it gives a mixture of violet and green light.

MERCURY, Hg, 200.

Sources. Found native, and as the sulphide (cinnabar) in Spain, California, Japan, and China. It was known to the ancients.

Preparation. The cinnabar is roasted; the sulphur is evolved as sulphur dioxide (SO_2), the metal volatilizes, and the vapor is condensed in earthen pipes.

Properties. The only metal liquid at the ordinary temperature. It freezes at -39°F . and C ., and boils at 675°F . (337°C .); when pure, does not tarnish in dry or moist air, but above 300°C . it absorbs oxygen and passes into the red oxide. It is largely employed in the processes for extracting silver and gold from their ores, and is used in medicine. It is a very lustrous and heavy metal, specific gravity, 13.56. Mercury has the power to dissolve metals, forming alloys which are known as amalgams. These are either soft or hard according to the quantity of mercury used. Mirrors are coated with an amalgam of tin. Dentists use extensively various amalgams for filling teeth. When mercury is triturated with a soft substance, it can be so finely divided as to lose all its lustre and appear as a bluish, or grey powder, which is used in medicine under the name, blue pill.

The vapor of the metal is 100 times as heavy as hydrogen, which is half as heavy as theory would require. Two series of mercury salts are known, corresponding to the copper salts, and called respectively, *mercurous* and *mercuric* salts.

Mercuric Oxide, HgO , *Red Precipitate*, can be obtained by keeping mercury for sometime at its boiling point, also by heating mercuric nitrate. It is a red or yellowish-red powder, decomposed at a dull red heat. It is an active base and dissolves very slightly in water. It is used in medicine. When caustic soda is added to a mercuric solution, mercuric hydrate, HgH_2O_2 , is thrown down as a yellow precipitate.

Mercuric Sulphate, HgSO_4 , is formed by boiling mercury with sulphuric acid. The reaction is



similar to that with copper, given on page 61. Mercuric sulphate is a white powder, which is decomposed by water, forming a yellow oxysulphate, $\text{HgSO}_4 + 2\text{HgO}$, called turpeth mineral.

Mercuric Nitrate, $\text{Hg}(\text{NO}_3)_2$, is generally seen in solution with excess of nitric acid, forming the *acid mercury nitrate*, used in medicine. The normal salt is used as a source of the oxide.

Mercuric Chloride, HgCl_2 , *Corrosive Sublimate*, is obtained by heating a mixture of mercuric sulphate and common salt.



The mixture is strongly heated; the mercuric chloride rises in vapor and condenses on a cool surface. This process is called *sublimation*. Corrosive sublimate is a heavy, white, crystalline powder, soluble in water and ether, and having an acid metallic taste. It is extremely poisonous; about five grains being a fatal dose. It forms with albumin, an insoluble precipitate.

Mercuric Iodide, HgI_2 , *Red Iodide*, is formed by mixing corrosive sublimate with potassium iodide.



Mercuric iodide is at first yellow, but rapidly changes to a brilliant scarlet. It is soluble, both in HgCl_2 and KI ; and the solutions are used as tests for ammonia and various organic principles.

Mercuric Sulphide, HgS , *Vermilion*, is found as a red mineral, *cinnabar*, an important ore of mercury. It can be prepared by passing hydrogen sulphide into mercury chloride,



Thus prepared it is black, but is converted into the red form by heating. It is used as bright red paint.

The *Mercurous Salts* are mostly of little importance.

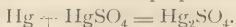
Mercurous Oxide, Hg_2O , is a black powder, easily decomposed, obtained by the action of alkaline hydrates upon calomel.

Mercurous Chloride, Hg_2Cl_2 , *Calomel*, is formed similarly to corrosive sublimate, substituting mercurous sulphate for mercuric.



The calomel is driven out by heat, and forms a sublimate. It is a white, heavy, tasteless powder, insoluble in water.

Mercurous Sulphate, Hg_2SO_4 , may be made by rubbing mercuric sulphate with another atomic proportion of mercury,



Flame Reaction. Mercury gives with the electric spark a mixture of green, blue and feeble yellow light.

SILVER, Ag, 108.

Sources. Occurs native in moderate abundance, also as sulphide, chloride and in other forms. It is generally present, in small amounts, in lead sulphide. Silver was known to the ancients.

Preparation. Silver is easily reduced to the metallic state, most of its compounds being decomposed by heat. When existing in small quantities in ores, it may be taken out by agitating the powdered material with mercury which dissolves the silver, (amalgamation) and this amalgam being drawn off and distilled, leaves the metallic silver. When compounds like sulphides or chlorides are reduced, it is sometimes necessary to add iron scraps to liberate the silver before adding the mercury.

Properties. Silver is a white metal of high lustre, easily worked into plates and wire, and the best conductor of heat and electricity known. Specific gravity, 10.5. It resists the action of oxygen and of caustic alkalies, but is attacked by sulphur and sulphides, and by nitric acid. Solutions of silver salts are reduced to the metallic state by electricity and by many forms of organic matter, especially when mixed with alkalies. The sensitiveness of its salts to light is the basis of photography. Silver melts at 1681°F . (916°C .) For coinage it is usually alloyed with copper. Silver differs from the other members of the lead group in acting as a monad, and in forming an aluin, and it may be regarded as a connecting link between this and the alkali group.

Silver Oxide, Ag_2O , cannot be formed directly, as, although oxygen is absorbed by melted silver, it is all given out as the metal cools. It is a black powder, usually made by heating

silver hydrate, AgHO , which latter is produced when silver nitrate is mixed with an alkali.



Silver Sulphate, Ag_2SO_4 , is sometimes used in analysis.

Silver Nitrate, AgNO_3 , *Lunar Caustic*, is easily made by dissolving the metal in nitric acid. The reaction is similar to that with copper.



Silver nitrate forms colorless crystals, very soluble in water, and blackened by light when mixed with organic matter. It fuses at 426°F . (219°C .), and is often cast in sticks for use as a caustic. The property of forming a black, difficultly soluble precipitate with organic matter, is utilized in the manufacture of hair dyes, and marking ink.

Silver Chloride, AgCl , is found as a mineral, and is easily formed artificially by adding any chloride to silver nitrate.



It forms a heavy white precipitate like curdled milk, turning violet in the light, especially if organic matter be present, forming a subchloride Ag_2Cl . It is insoluble in most acids, but dissolves freely in ammonia, and in sodium thiosulphate.

A number of other silver compounds have been described, but have but little importance.

Silver Salts in Photography. The action of light upon silver compounds is mostly an operation of reduction, sometimes to the metallic condition. The presence of organic matter aids the change. In some of the compounds the effect of the light is not attended with any change of color.

Extraction of Silver from Residues. As silver is a valuable metal, it is profitable to extract it from wastes and refuse material. Two methods are used, (a) wet method; the silver is dissolved, if not already in solution, by nitric acid, and common salt added, until all the silver is thrown down as chloride. This precipitate is washed with clean water, and mixed with zinc and dilute sulphuric acid. The chloride begins at once to reduce to metallic silver. The operation is complete when the mass has become of a uniform dark grey color. The metallic silver is washed thoroughly with water, and fused if necessary: (b) dry method. Silver chloride is prepared as before, washed thoroughly, dried and mixed with about twice its weight of dry sodium carbonate, and fused in a crucible at a full red heat for some time. The crucible is broken when cold, and a button of silver will be found at the bottom.

Flame Reaction. Silver gives no color to flame; with the electric spark it gives a green color mixed with a feeble blue and violet.

THALLIUM, Tl, 204, was discovered by Crookes in 1861. It exists in some varieties of iron pyrites, and in some mineral waters. It acts as a monad and triad. The metal resembles lead in many points; specific gravity, 11.8, but its compounds are somewhat like those of the alkali metals, and somewhat like those of silver. It forms two oxides, Tl_2O and Tl_2O_3 , the former being very soluble in water; two chlorides, $TlCl$ insoluble in water, and $TlCl_3$ soluble in water, and other salts of similar relations. Thallium compounds give a pure green color to flame.

GOLD, Au, 196.7.

Sources. Occurs in the metallic state, often in veins in quartz, often in small grains in sand and gravel; sometimes alloyed with silver, copper or other metals. Gold was known to the ancients.

Preparation. It is often extracted by washing which carries off the lighter sand and dirt and leaves the gold. Quartz rock is first ground. It is sometimes extracted by amalgamation, as described under silver.

Properties. Pure gold is a bright yellow heavy metal, (specific gravity, 19.4), very soft and more capable than any other metal of being worked into thin plates or wire, and an excellent conductor of heat and electricity. It melts at $1900^{\circ} F.$ ($1036^{\circ} C.$) Gold is unaffected by air, water or sulphur or by ordinary acids, even at high temperatures. Its compounds are reduced to the metallic state by heat, and by reducing agents in the cold. Chlorine, and a mixture of nitric and hydrochloric acid, (which contains free chlorine) dissolve it forming chloride. In the pure condition, it is very soft and can be welded in the cold by pressure. Gold foil is prepared in this form for dentists' use. For jewellery, coin and other articles subjected to wear, the pure metal is too soft and is alloyed with copper and silver. The proportion of alloy is indicated by *carats*, pure gold being 24 carats, 18 carat gold being 18 parts gold and four parts alloy. Copper makes a red gold, silver, a green gold. In its chemical relations gold does not resemble the metals, and it is probable that it would be placed among the metalloids, if its physical properties of lustre, weight and insolubility were not so distinctly metallic. Two sets of compounds are known, aurous and auric, in

which the metal is respectively a monad and triad. The oxides are not bases; one appears to be an anhydride.

Au_2O	Aurous oxide.
Au_2O_3	Auric “
AuCl	Aurous chloride.
AuCl_3	Auric chloride.

Auric Chloride is produced when gold is dissolved in nitromuriatic acid. By adding to the liquid free from excess of acid, some ferrous sulphate, the gold is thrown down in the metallic state as a brown powder, looking like mud. A mixture of stannous and stannic chlorides (q. v.) produces with gold chloride a purple precipitate called *purple of Cassius*, which is used for coloring glass and porcelain.

Group 10. The Iron Group. Includes iron, manganese, nickel, cobalt and partly, chromium. They are not found in the metallic state, except in small quantity. They form two sets of compounds, acting in one set as a dyad, in the other as double tetrads. Several of them are capable of forming well marked anhydrides. Their atomic weights are all nearly the same.

NICKEL, Ni, 59.

Sources. Occurs principally in union with arsenic and sulphur; also in meteoric iron as an alloy. Discovered by Cronstedt in 1751.

Preparation. Nickel is reduced by roasting and reduction with charcoal.

Properties. It is a hard white metal, specific gravity 8.8, fusing at a high temperature, and resisting the action of air at common temperatures. It is one of the few metals which can acquire permanent magnetism. Solution of nickel can be decomposed by an electric current; and nickel plating is performed in this way. An alloy of copper, zinc and nickel is called German silver, and some other alloys have been used in coins.

Nickel Monoxide, NiO , is obtained by heating the nitrate, or carbonate. A hydrate, NiH_2O_2 , is obtained in the usual manner. It is green, and soluble in ammonia. Its compounds with acid are mostly bright green.

Nickel Sesquioxide, Ni_2O_3 , is also known, but does not appear to form salts.

Nickel Sulphate, NiSO_4 , is the most important salt. It usually crystallizes with 7 molecules of water, and can easily form double salts with potassium and ammonium sulphates.

Flame Reactions. Nickel compounds give with the electric spark a mixture of colors.

COBALT, Co, 59, discovered by Brandt in 1733, is found associated with nickel, which it closely resembles in properties and chemical relations. Its compounds are mostly red or blue. The metal is hard white difficult to fuse and magnetic; specific gravity, 8.7. The oxides, sulphates, carbonates, etc., resemble in composition those of nickel. It produces with the electric spark a mixture of colors.

IRON, Fe, 56.

Sources. Occurs very abundantly as oxide, sulphide and carbonate; some of its compounds occur in small quantities in many rocks and soils, and hence it is dissolved by natural waters and taken up by plants; from these sources it gets into the animal body. It is an essential constituent of the blood of the higher animals. Iron sometimes occurs native. It was known to the ancients.

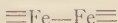
Preparation and Properties. The preparation of iron is a simple process of reduction. A fine grade, Quevenne's iron, for use in medicine, is made by passing hydrogen over sesquioxide. On the large scale, iron ore, which generally consists of an oxide, or carbonate, mixed with clay, sand and other minerals, is heated in blast furnaces with coal and limestone. The limestone makes a fusible calcium silicate, *slag*; the coal takes the oxygen away from the iron. The melted metal is then run out into thick bars, forming pig, or cast-iron, the most impure form; containing four or more per cent. of carbon, also phosphorus, sulphur, silicon and other bodies. This crude material is worked in a furnace in a current of air, by which the impurities, especially the carbon, are burnt out, and the iron becomes less fusible and much tougher, constituting *wrought iron*. *Steel* is iron containing about one-half per cent. of carbon; it is therefore intermediate in composition. Steel was formerly made by heating wrought iron with charcoal; it is now made directly from cast iron by burning out a certain percentage of the carbon, (Bessemer process.) Pure iron is very soft, but as found in commerce the various impurities give hardness and other special qualities to the iron; carbon gives hardness and fusibility, phosphorus and sulphur give fusibility and great brittleness, and are very objectionable. Cast iron

melts about 3000° F. (1649°C.) Iron is strongly magnetic and not much affected by dry air, but is oxidized by moist air and easily dissolved by acids. It forms two series of salts; *ferrous*, in which the metal is a dyad, and *ferric*, in which it is apparently a triad, but the formulæ of the ferric compounds are generally doubled, and the metal is regarded as a double tetrad acting as a hexad, (see pages 16, 17, 19), thus :

Dyad iron,



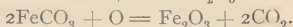
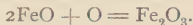
Hexad iron



Ferrous salts are generally green; ferric salts brown or red.

Conversion of one Series of Iron Salts into the Other.

Ferrous salts are converted into ferric by oxidizing agents. As the ferric salts are written with double formulæ, the reaction will always require two molecules of ferrous for one of ferric, thus :

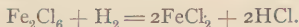


As nitric acid gives up three atoms of oxygen as an oxidizing agent, (page 44), it follows that in writing reactions, in which it is used, six molecules of the ferrous salt must be taken.

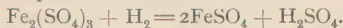
To make a normal ferric salt, one additional molecule of the acid from which the salt was formed must be added for every two molecules of the ferrous salt. For instance, in making ferric chloride, if we take $6\text{FeCl}_2 + 2\text{HNO}_3$, we will get an oxychloride, or oxynitrate. The complete reaction is :



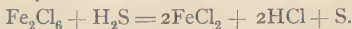
Ferrous salts are formed from ferric by the action of reducing agents, especially nascent hydrogen or sulphurous acid. The nascent hydrogen is usually obtained by adding to the ferric salt a mixture of zinc and sulphuric acid. With ferric chloride the reaction would be,



With ferric sulphate we would have,



Hydrogen sulphide will also reduce ferric salts.

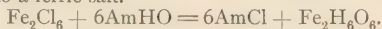


Ferrous Oxide, FeO , is difficult to obtain. Ferrous hydrate FeH_2O_2 is formed as a white precipitate by the action of caustic alkali upon a ferrous salt. It immediately begins to change by absorbing oxygen, and becomes ferric oxide.

Ferric Oxide, Fe_2O_3 , *Red Oxide*, *Sesquioxidé*, occurs frequently in small amounts in many minerals, and also as iron

ore, called red hæmatite, or specular iron. It may be prepared artificially by heating ferrous sulphate, ($2\text{FeSO}_4 = \text{Fe}_2\text{O}_3 + \text{SO}_2 + \text{SO}_3$) and is the residue obtained in making Nordhausen sulphuric acid. It is a soft red powder, difficult to dissolve in acid. The finer grades constitute *rouge*, the coarser, *venetian red* and *crocus*, are used for paints.

Ferric Hydrate, $\text{Fe}_2\text{H}_6\text{O}_6$, is easily formed by adding caustic alkali to a ferric salt.



Ferric hydrate is a soft, brown mass insoluble in water, but dissolving easily in acid. Its chief importance is an antidote to arsenic, for which use it must be freshly prepared. Ordinary iron rust consists of impure ferric hydrate, and it also occurs in an impure condition as brown hæmatite, a very important ore of iron.

Magnetic Iron Oxide, $\text{FeO}, \text{Fe}_2\text{O}_3$, a union of the two oxides is found as a finely crystallized mineral, and valuable ore of iron. It can retain magnetism, and is occasionally found in a magnetized condition, constituting *loadstone*.

Ferrous Sulphide, FeS , made by fusing iron with sulphur, is a dark slag-like mass, used as a source of hydrogen sulphide.

Iron Disulphide, FeS_2 , *Iron Pyrites*, is abundant as a mineral crystallized in brass-colored cubes, often mistaken for gold, and hence called fool's gold. It is of no use as an iron-ore on account of the sulphur, but is used as a source of sulphuric acid, and some sulphates.

Ferrous Carbonate is a mineral, and valuable iron-ore, and exists in many mineral (chalybeate) waters. It is produced by mixing ferrous sulphate with sodium carbonate.

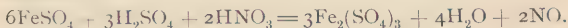


In this form, however, and also as dissolved in water it is prone to oxidation, passing into the condition of ferric hydrate, which forms a red deposit. This oxidation is hindered by sugar, and a ferrous carbonate, known as Vallet's mass, is prepared for medical use by precipitating it in contact with sugar. The natural carbonate is permanent in the air.

Ferrous Sulphate, FeSO_4 , *Green Vitriol*, *Copperas*, is formed by dissolving iron in sulphuric acid, or by oxidizing iron pyrites. It forms clear green crystals containing $\text{FeSO}_4 + 7\text{H}_2\text{O}$, easily soluble in water, the solution being liable to oxidation. It is used in the manufacture of ink and in dyeing. The crystals effloresce on exposure to air. Ferrous sulphate is a reducing agent, and is called a disinfectant, but merely acts as a deodorizer by absorbing some sulphur compounds. It forms double salts with ammonium and potassium sulphates,

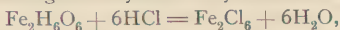
Ferric Carbonate has not been obtained.

Ferric Sulphate, $\text{Fe}_2(\text{SO}_4)_3$, is made by heating together a mixture of ferrous sulphate, sulphuric acid and nitric acid. The reaction is



If one-half the above quantity of sulphuric acid be used an oxysulphate, $5\text{Fe}_2(\text{SO}_3)_3 + \text{Fe}_2\text{O}_3$ is formed, called *Monsel's Salt*, or when dissolved *Monsel's solution* and used as a styptic.

Ferric Chloride, Fe_2Cl_6 , is generally seen as an alcoholic solution, called *muriated tincture of iron*. Ferric chloride is made either by dissolving ferric hydrate in hydrochloric acid,



or by boiling ferrous chloride with nitric and hydrochloric acid,



When the solution is evaporated a red crystalline mass of $\text{Fe}_2\text{Cl}_6 + 6\text{H}_2\text{O}$ is formed, which is decomposed by heat. The anhydrous salt is yellow, and extremely deliquescent. The solution dissolves ferric hydrate, forming a ferric oxychloride.

Iron nitrates, iodides and bromides do not possess special interest.

Flame Reaction. Iron gives no color to gas flames; with the electric spark it gives a mixture of many tints.

MANGANESE, Mn, 55.

Sources. Exists principally as oxide, also a sulphide, carbonate and silicate. Discovered by Gahn in 1780.

Preparation and Properties. By reducing the oxide or carbonate with charcoal, or by electrical decomposition of the chloride. It is a greyish-white metal, brittle and hard, specific gravity between 7 and 8. It forms two series of salts parallel to those of iron.

Manganous Oxide, MnO , is obtained by igniting the carbonate; the *hydrate*, MnH_2O_2 , is obtained by adding an alkali to a manganous salt. Both substances absorb oxygen rapidly.

Manganous Chloride, MnCl_2 , may be obtained from the residue left after the preparation of chlorine. It forms pink crystals, deliquescent, and soluble in water.

Manganous Sulphate, MnSO_4 , is made by dissolving the oxide, or carbonate in sulphuric acid. It is a rose-colored salt, soluble in water, and used in dyeing. It crystallizes with $7\text{H}_2\text{O}$.

Manganese Dioxide, MnO_2 , *Black Oxide*. This is rather an abundant mineral, occurring crystallized or in masses. It is extensively used as an oxidizing agent, and in the manufacture of chlorine. Ordinarily it is in black masses, or powder. It conducts electricity, and dissolves in acids with evolution of oxygen. When heated to redness, it is converted into an intermediate oxide, Mn_3O_4 .

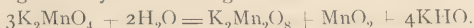
Manganic Oxide, Mn_2O_3 , is found as a mineral. It is a weak base.

MANGANESE ACIDS. Manganese is capable of forming two anhydrides, which are not known in the free state, but of which some of the salts are common. The acids are



The former has not been obtained.

Manganates. These are formed by fusing manganese dioxide with caustic alkali and potassium nitrate, or chlorate. In this way potassium manganate K_2MnO_4 is formed as a green crystalline mass. When dissolved in water the manganates turn into permanganates by the following reaction :



The change of composition is indicated by a change of color from green to red, for which reason the potassium manganate was called *chameleon mineral*.

Potassium Permanganate, $\text{K}_2\text{Mn}_2\text{O}_8$, is now made in large amount as an oxidizing and deodorizing agent. The solution slowly decomposes when exposed to the air, and is supposed to give off ozone, and thus act as a disinfectant. It is decomposed by organic matters, by sulphites and sulphides and reducing agents generally, becoming converted into a colorless solution. It can, therefore, be employed not only to destroy organic matter, but also as a measure of the amount present.

Flame Reaction. Manganese compounds give to gas flame a mixture of yellow and green light; with the electric spark a number of tints, orange, yellow and blue are seen.

CHROMIUM, Cr, 52.5.

Sources. Occurs principally as an oxide in combination with iron oxide, constituting chrome iron ore, $\text{FeO}, \text{Cr}_2\text{O}_3$; also as lead chromate, PbCrO_4 . Discovered by Vanquelin in 1797.

Preparation and Properties. By heating the oxide with charcoal, chromium is obtained as a hard crystalline metal not easily oxidized or dissolved. It forms two sets of salts analogous to those of iron, and also a marked anhydride, which forms salts isomorphous with the sulphates. By this fact chromium is partly related to the non-metals. The compounds in which it acts as a positive metal are of very little importance; they agree mostly with the corresponding iron and manganese compounds. The chromous salts are unstable. Almost all the chromium compounds are high colored.

Chromic Oxide, Chromium Sesquioxide, Cr_2O_3 , Chrome Green, obtained by decomposing some of the chromates (q. v.) is a bright green powder, used as a paint.

Chromic Anhydride, CrO_3 , is easily obtained by the action of acids upon the chromates. It forms bright red crystals, very deliquescent, soluble in water and having powerful oxidizing properties.

Chromates are formed by heating chrome iron ore with alkali and nitre. The most common salts are those of potassium.

Potassium Chromate, K_2CrO_4 , forms lemon yellow crystals soluble in water.

Potassium Anhydrochromate, $\text{K}_2\text{CrO}_4\text{CrO}_3$, commonly known as *bichromate*, is in large, bright red crystals which are soluble in water. It is extensively used in dyeing, and as a source of various colors.

Lead Chromate, Chrome Yellow, PbCrO_4 , is easily formed by adding a soluble chromate to a lead salt.

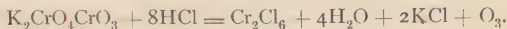


It is bright yellow, and insoluble in water. An oxychromate $\text{PbCrO}_4 + \text{PbO}$, is known as a scarlet pigment.

Chromates are decomposed when heated with organic matter, especially in the presence of an acid. The change of composition is generally indicated by a change of color from the yellow, or red of the chromate to the green of the chromic salt.

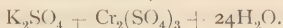
Exp. Add to a dilute solution of potassium anhydrochromate, a few drops of alcohol and some hydrochloric acid, and boil the mixture for a few moments. Vapors of aldehyde (q. v) will be given off, and the red color of the liquid will change to green, chromic chloride being formed.

The reaction is :



The O_3 oxidizes the alcohol to aldehyde.

A mixture of potassium anhydrochromate and sulphuric acid is used as an oxidizing agent in galvanic batteries. The chromic acid becomes reduced to sesquioxide, and forms chromic sulphate; the liquid turns green, and afterwards deposits dark ruby-red crystals of chrome-alum,



Flame Reaction. Chromium communicates no color to flame; with the electric spark it gives a mixture of green and dark blue.

URANIUM, U, 240, was discovered by Klaproth in 1789. It is rather rare, occurring as oxide and phosphate. It is a greyish metal not oxidized by air, or water, but dissolving in acid. It forms two classes of compounds, uranous (tetrad) and uranic, (hexad). Uranic oxide UO_3 can act as an anhydride, forming uranates. Uranium compounds are generally yellow or yellowish-green and are *fluorescent*, that is, shine brilliantly under the influence of rays of light to which the eye is ordinarily not sensitive. Glass colored yellowish with uranium oxide is much used in optical experiments.

VANADIUM, V, 51.2, discovered by Del Reo in 1801, is a rare body, found chiefly in combination with iron and lead. It belongs with phosphorus and nitrogen, and forms four oxides, VO , V_2O_3 , VO_2 , V_2O_5 analogous to those of nitrogen.

Vanadic Anhydride, V_2O_5 , forms of salts, called vanadates. Lead vanadate is found as a mineral. It yields compounds analogous to metaphosphoric acid, and also forms salts with some of the strong acids. Vanadium has acquired some importance from the possibility of making from it a good indelible ink, but the rarity of its compounds has interfered with their use.

OSMIUM, Os, 199, discovered by Tennant in 1803, is associated with platinum. It is strongly metallic in many of its properties, having high lustre, specific gravity, 21.4, very difficult to fuse and forms an amalgam with mercury. Five oxides are known, OsO , Os_2O_3 , OsO_2 , OsO_3 , OsO_4 . The monoxide and sesquioxide form an imperfect series of salts; the trioxide and tetroxide form salts with bases.

MOLYBDENUM, Mo, 96, discovered by Bergman in 1781, occurs chiefly as disulphide, and as lead molybdate, both rather rare minerals. It forms three oxides, MoO , MoO_2 , MoO_3 . The last, *molybdic anhydride* forms salts called molybdates. Ammonium molybdate Am_2MoO_4 is used as a test for phosphoric acid, with which it forms a highly insoluble yellow precipitate.

BISMUTH, Bi, 210.

Sources. Commonly found native; also as oxide and sulphide. It was known to the earlier chemists.

Preparation and Properties. The metal is easily extracted by fusion in iron cylinders. It is hard, brittle, reddish-white, and distinctly crystalline. It fuses at 507° F. (264° C.), expanding when it solidifies. Its chemical relations are somewhat like those of antimony, but it forms well marked salts and must, therefore, be regarded as a metal. It is not much affected by the air. Nitric acid dissolves it easily. The bismuth oxides are Bi_2O_2 , Bi_2O_3 , Bi_2O_4 , Bi_2O_5 .

Bismuth Sesquioxide, Bi_2O_3 , the only important oxide is obtained as a yellowish powder by burning bismuth in the air, or by heating the carbonate or nitrate. It acts as a base.

Bismuth Nitrate, $\text{Bi}(\text{NO}_3)_3$, made by dissolving the metal in nitric acid, is a soluble, white, crystalline mass. When added to a large volume of water, a white precipitate of bismuth oxynitrate of irregular composition, but generally $\text{Bi}(\text{NO}_3)_3 + \text{Bi}_2\text{O}_3$, is thrown down. This powder ordinarily called *bismuth subnitrate* is used in medicine, and sometimes as a cosmetic. When it is boiled with caustic soda and a solution of glucose, a heavy black powder of metallic bismuth is formed. This reaction constitutes Boettger's test for sugar.

Bismuth Chloride, BiCl_3 , is decomposed by water in a manner similar to the nitrate.

Group 11. The Tin Group. These elements all have the physical properties of metals; lustre, weight and insolubility, but their chemical properties bring them near to the non-metals, especially to silicon. They are, therefore, classed as metalloids. They form oxides which in the presence of strong acids have feeble basic properties, and agree with the non-metals in forming a well marked anhydride.

TIN, Sn, 118.

Sources. Occurs principally as dioxide, called tin stone. The metal was known to the ancients.

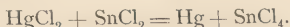
Preparation and Properties. The ore is roasted and reduced with charcoal. The metal is white, soft and easily beaten

into foil, but is not tough; specific gravity, 7.28. It fuses at 442° F. (228° C.) It resists very well the action of the air, and of cold acids. Nitric acid forms an insoluble dioxide. Tin forms several valuable alloys: pewter, gun-metal, type-metal, bronze and solder, elsewhere described. Speculum metal, used for metal mirrors is an alloy of copper and tin; glass mirrors are coated with an amalgam of tin. Tin plate is iron coated with tin by dipping it into a bath of the melted metal.

Two series of tin salts are known, *stannous*, dyad, and *stannic*, tetrad.

Stannous Oxide, SnO , is a feeble base, but is not important.

Stannous Chloride, SnCl_2 , is formed by dissolving tin in hydrochloric acid. The solution deposits white crystals containing $2\text{H}_2\text{O}$. They dissolve in water, but are generally quickly decomposed into an oxychloride, which precipitates. Stannous chloride is an unsaturated molecule, and tends to take up chlorine or oxygen, for which reason it is used as a reducing agent. When mixed with mercuric chloride, metallic mercury is set free, and stannic chloride formed.



Stannous chloride is used by the dyer as a mordant, under the name of tin crystals.

Stannous Sulphide, SnS , is found as a mineral, and is obtained artificially by the action of hydrogen sulphide upon stannous chloride.



Stannic Oxide, SnO_2 , *Stannic Anhydride*, is found as a mineral; also produced by burning tin in the air, by oxidizing it with nitric acid and by adding an alkali to stannic chloride. Like silicic anhydride it forms different acids according to the method of production. When tin is heated with nitric acid, a white powder is obtained which forms complicated salts called *metastannates*. When an alkali is added to the chloride, an acid is formed, having the composition H_2SnO_3 , forming salts called *stannates*. This latter acid is truly analogous to metasilicic, H_2SiO_3 , and should have been called metastannic.

Stannic Chloride, *Tin Tetrachloride*, SnCl_4 , *Libavius' fuming liquor*. This body is largely used by dyers under the name of *nitro-muriate of tin*, being made by the action of a mixture of nitric and muriatic acid on metallic tin. It is a colorless fuming liquid, boiling at 239.5° F. (115.3° C.)

Stannic Sulphide, SnS_2 , *Mosaic Gold*, is made by passing hydrogen sulphide into stannic chloride. It is a bronze colored powder used in printing and coloring.

TITANIUM, Ti, 48, discovered by Gregor in 1789, exists as titanic anhydride, TiO_2 , and also in some iron ores.

Tantalum, Ta, 182, discovered by Hatchett in 1801, and *Niobium*, Ni, 94, are found in some rare minerals.

TUNGSTEN, W. 184, discovered by Bergman in 1783, exists as a manganese-iron tungstate, called wolfram, and as calcium tungstate, (scheelite.) The only important compound of the metal is tungstic acid, H_2WO_3 . Sodium tungstate has been employed for rendering dress goods fire-proof.

Zirconium, Zr, 90, discovered by Klaproth in 1789, probably belongs in the aluminium group. It exists as oxide in the rare mineral *zircon*.

Group 12. The Platinum Group. This includes several metals, which are nearly always found in the metallic condition, and generally alloyed with one another. They probably do not all belong to one class, but their rarity has prevented the extensive study of their compounds.

PLATINUM, Pt, 196.7.

Sources. Occurs native, alloyed with the metals of its class, and also with gold, silver, etc. Discovered by Wood in 1741.

Preparation and Properties. The ore is purified by solution in *aqua regia*, and precipitation of the platinum as ammonio-platinum chloride, $2\text{AmCl} + \text{PtCl}_4$. This precipitate is decomposed by heat, leaving the metal in a spongy state. This is compressed into a small bulk and hammered while red-hot. It can also be fused in the flame produced by burning a mixture of oxygen and hydrogen.

Platinum is a hard white metal, very heavy; specific gravity, 21.5, it fuses only at a very high temperature. It resists perfectly the action of the air, and of most chemical agents, and for this reason is largely used in chemical operations. It dissolves in hot *aqua regia*, forming platinum tetrachloride, PtCl_4 . A number of the common metals, *e. g.* zinc, tin, lead, and silver, when melted with platinum, dissolve it, forming fusible alloys. Platinum forms two series of compounds; its oxides are only feebly basic.

Platinum Tetrachloride, or Platinic Chloride, PtCl_4 , is a red or brown deliquescent mass, obtained by dissolving the metal in

aqua regia, and evaporating the solution to dryness. It forms yellow granular precipitates with potassium and ammonium salts, but not with those of sodium, and is of great use in analysis for the separation of potash and soda.

PALLADIUM, Pd, 106.5, discovered by Wollaston in 1803, exists associated with platinum and gold. It is a hard white metal, specific gravity, 11.6, and is not easily oxidized. Hammered palladium absorbs 640 times its volume of hydrogen, forming an alloy. Two series of palladium compounds are known in which the metal is respectively dyad and tetrad.

RUTHENIUM, Ru, 104.4, discovered by Claus in 1846, is found with platinum; also as a sulphide. It is a hard brittle metal, difficult to fuse and to oxidize; specific gravity, 11.4. It acts as a dyad and tetrad.

RHODIUM, Rh, 104.3, discovered by Wollaston in 1803, exists in platinum ores. It is a hard brittle metal, melting only at a high temperature, and then oxidizing. Specific gravity, 12.1. The atomicity of rhodium seems to resemble that of iron; it forms a monoxide and sesquioxide.

• IRIDIUM, Ir, 198, discovered by Tennant in 1803, is found with platinum and osmium. It is a hard, white, not very tough metal, which is difficult to fuse and to dissolve. It forms like rhodium, compounds which recall those of iron. We have iridous chloride, IrCl_2 , and iridic chloride, Ir_2Cl_6 .

ORGANIC CHEMISTRY.

Organic Bodies were at first understood to be those that exist only in living structures. The progress of chemistry made known many substances which could be produced by artificial means from the true organic bodies, and thus organic chemistry came to include not only the constituents of animals and plants, but the derivatives from them. A still further advance was made when some of these substances were produced directly from their elements without the intervention of life, thus showing that inorganic and organic bodies were not essentially different. Although, therefore, the essential distinction between the two branches of the science has faded away, it is still convenient to describe under one section a great many bodies which *can* be made artificially, but which are generally made from living tissue. These so-called organic bodies are distinguished by a few important peculiarities. The greater number of them contain carbon, hydrogen and oxygen. Some contain only carbon and hydrogen; others contain nitrogen. The extensive variety of compounds is due to variations in proportion and arrangement of atoms. The organized tissues of plants and animals—muscle, brain, skin, seed, leaf, pollen, etc.—are generally complex in character and often contain other elements, especially sulphur and phosphorus, in addition to those mentioned above. Many organic bodies are subject to change and, as a rule, the more complex the substance, the more easily is it decomposed. The number of organic bodies has been greatly increased by the action of certain chemical agents upon the substances naturally occurring in plants and animals.

(a) *Action of Heat.* Heat usually decomposes organic substances leaving a residue of carbon, and producing a mixture of bodies—solid, liquid and gaseous.

(b) *Action of Oxygen.* Oxygen when acting at low temperature generally forms acids, either by direct union or by removing the hydrogen and taking its place.

(c) *Action of Nitric Acid.* This sometimes produces a simple oxidation. Very often it removes hydrogen and adds NO_2 .

(d) *Action of Chlorine.* Chlorine generally removes hydrogen and takes its place atom for atom. Bromine and iodine have the same action.

This substituting action of nitric acid and halogens gives rise to a large series of compounds, organic in general relations, but

not existing in any living tissue. The substitution of NO_2 gives us a series of nitro-compounds, and takes place in the proportion of one molecule of NO_2 for each atom of H.

(e) *Action of Dehydrating Agents.* Dehydrating agents (bodies which have strong affinity for water) generally remove both hydrogen and oxygen from the organic compound in the proportion of two atoms of H to one of O, (H_2O .)

(f) *So-called Natural Changes.* These are FERMENTATION, PUTREFACTION and DECAY. Fermentation is a change produced in an organic body by the action of a decomposing nitrogenous body, whereby it is reduced to a simpler form. Putrefaction is a change taking place especially in nitrogenous bodies, merely under the influence of the ordinary conditions. Decay is a modified putrefaction in which oxidation occurs.

The tissues of plants and animals, and sometimes the substances derived from them, are mixtures of two or more distinct compounds. The separation of these bodies from each other is called PROXIMATE ANALYSIS, and the bodies so separated are called proximate principles. The determination of the proportions of elements in any body, is called ULTIMATE ANALYSIS.

The complexity of organic bodies is such that two bodies may have the same composition, but owing to a different *arrangement* of the atoms, be different bodies. This is called ISOMERISM.

A formula which expresses simply the proportion between the atoms of elements present in a compound is called an EMPIRICAL FORMULA. A formula which in addition expresses the arrangement and exact number of atoms is called a RATIONAL FORMULA.

Organic chemistry presents us with a number of compound bodies which have the characters of elements in so far that they form distinct oxides, chlorides, etc., and replace the hydrogen of acids. They are, therefore, quasi-metals analogous to ammonium (page 103); they are generally called ORGANIC RADICLES, (often incorrectly spelled *radicals*.) The most important of them are compounds of carbon and hydrogen. Their number is greatly increased by the power which the carbon possesses (at least this is the accepted theory) of combining with itself, forming duplicated atoms, or as they are called *carbon skeletons*. In this manner we will have

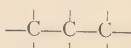
Carbon, tetrad.



Dicarbon, hexad.



Tricarbon, octad.



It will be noticed that the addition of each carbon atom adds two degrees to the atomicity, to saturate which two atoms of hydrogen will be required. It follows, therefore, that if we add

CH_2 to any radicle, we do not change its atomicity. In CH_3 for instance, we have a monad radicle of which the graphic formula may be given $-\text{C}\equiv\text{H}_3$; if to this we add $=\text{C}=\text{H}_2$, we will get $\text{H}_2=\text{C}-\text{C}\equiv\text{H}_3$, in which one bond is still unsatisfied.

In accordance with this principle we have a number of series, the members of which differ by CH_2 , and possess the atomicity of the lowest member. Such a series differing by CH_2 is said to be HOMOLOGOUS.

Analysis of Organic Bodies. The best general test for an organic substance is the action of heat which usually causes decomposition with evolution of smoky, strong smelling vapors and leaves a residue of carbon which can be burned off by heating strongly in the air. Sulphuric acid also produces a characteristic blackening, due to liberation of carbon. The presence of nitrogen is usually indicated by a very disagreeable odor; more surely by heating the body with an alkali by which ammonia is formed. The accurate analysis of organic bodies is performed by burning them completely in a current of oxygen, collecting and weighing the carbon dioxide, water and nitrogen, which are thus produced. This process gives the *ultimate* composition. To separate the various organic substances existing in a mixture and thus make a *proximate* analysis, is a matter entirely dependent on the nature of the substances present; it is sometimes very easy, and in other cases so difficult as to be practically impossible. Solids are generally separated by differences of solubility in water, alcohol, ether and other solvents; liquids are separated by *fractional distillation* in which advantage is taken of different boiling points. Gases are separated by different absorbents. The principle of these methods may be illustrated by making a mixture of starch, sugar, rosin and some volatile oil. By gently heating the mixture the oil may be distilled off; cold water will take up the sugar, and hot water the starch; alcohol will then dissolve the rosin.

Classification of Organic Bodies. Two methods of classification are in use: the older system was to arrange bodies according to similarity of properties or similarity in methods of production; the more recent system seeks to arrange bodies according to similarity of composition and chemical relation. The latter method is, of course, the more scientific, but it is doubtful whether it is the best for the beginner. It is preferable to make, in an elementary work, a combination of methods, and describe together those bodies of generally similar character. We make under this arrangement the following classification:

Hydrocarbons.

Bodies containing hydrocarbons as radicles, { alcohols,
ether, etc.

Sugars and starches.

Oils and fats, including fat-acids.

Organic acids, not fatty.

Resins and coloring matters.

Alkaloids and active principles of plants.

Hydrocarbons. Many of these are produced by the destructive distillation of organic substances, especially coal and wood. They are arranged in a number of homologous series of which only the more important will be given. The fusing and boiling points gradually rise as the number of carbon atoms increases, so that if the lower members are gases, those in the middle of the series will be liquids, the higher members will be solids. In each series the atomicity is determined by that of the first member, because, as explained on page 137, the addition of CH_2 does not affect the saturation.

PARAFFINS. This series begins with marsh gas CH_4 . Being saturated bodies they are comparatively indifferent to chemical reagents.

Marsh Gas or Methane is produced by distillation of wood and coal or by decay of vegetable matter. It is next to hydrogen the lightest body known. It exists in coal gas.

Common Paraffin exists in petroleum and in coal-tar. It is a mixture of several of the higher members of the series. It is a white waxy solid, easily fusible, little acted on by acids or alkalis and used for a protecting coating in chemical apparatus. Cosmoline and its imitations are also in part soft paraffins..

METHYLS. These begin with methyl CH_3 and are a series of monad radicles forming the monatomic alcohols.

Some of the important members of the series are here given.

CH_3	Methyl.	C_3H_7	Propyl.
C_2H_5	Ethyl.	C_4H_9	Butyl.
	C_5H_{11}	Amyl.	

These form compounds analogous in structure, but not in appearance or general properties, to those of the alkali metals. The compounds are very important. We have

1. Normal oxides, called **SIMPLE ETHERS.**

$(\text{CH}_3)_2\text{O}$ Methyl ether analogous to Na_2O , sodium oxide.

$(\text{C}_2\text{H}_5)_2\text{O}$ Ethyl ether, " " " "

2. Compounds with halogens, sometimes called **ETHERS.**

$(\text{CH}_3)\text{Cl}$ Methyl chloride analogous to NaCl , sodium chloride.

$(\text{C}_5\text{H}_{11})\text{Cl}$ Amyl " " " "

3. Compounds analogous to salts, called COMPOUND ETHERS.
 $(\text{CH}_3)_2\text{SO}_4$ Methyl sulphate analogous to Na_2SO_4 , sodium sulphate
 $(\text{C}_5\text{H}_{11})\text{NO}_3$ Amyl nitrate " NaNO_3 , sodium nitrate.

5. Compounds analogous to the acid salts called VINIC ACIDS.

$(\text{C}_2\text{H}_5)\text{HSO}_4$ Sulphethylic acid, analogous to KHSO_4 .
 $(\text{C}_5\text{H}_{11})\text{HSO}_4$ " amylic " " " " "

6. Compounds analogous to the hydrates, called ALCOHOLS.

$(\text{C}_2\text{H}_5)\text{HO}$ Ethyl alcohol, analogous to KHO .

$(\text{C}_5\text{H}_{11})\text{HO}$ Amyl alcohol, " "

7. Compounds containing two different radicles analogous to the mixed salts and called MIXED ETHERS.

$(\text{CH}_3)(\text{C}_2\text{H}_5)\text{O}$ Methyl-ethyl ether.

OLEFINS. These begin with C_2H_4 , olefiant gas, ethylene, which is contained in coal-gas and can be made pure by the action of dehydrating agents on alcohol. The olefines are dyads, and form compounds which agree in structure with those of the common metals. $(\text{C}_2\text{H}_4)\text{Cl}_2$, ethylene dichloride corresponds to CuCl_2 , copper dichloride.

The hydrates constitute the diatomic alcohols or glycols: $(\text{C}_2\text{H}_4)\text{H}_2\text{O}_2$, called ethylene glycol, is analogous to CuH_2O_2 , cupric hydrate.

METHENYL SERIES. These are triad radicles beginning with CH , methenyl, which forms $(\text{CH})\text{Cl}_3$ analogous to BiCl_3 .

The series may be regarded as forming the triatomic alcohols or glycerins.

$(\text{C}_3\text{H}_5)\text{H}_3\text{O}_3$, glycerin is analogous to BiH_3O_3 , bismuth hydrate.

TERPENES. These are principally natural products constituting the volatile or essential oils. C_5H_8 , called quintone, is the lowest member of the series known.

Oil of Turpentine, $\text{C}_{10}\text{H}_{16}$, is obtained from turpentine, an exudation from pine trees, and consisting of resin and volatile oil. On being distilled, the volatile oil is collected in the receiver; the resin remaining constitutes common rosin. Oil, or spirits of turpentine is a thin colorless liquid of peculiar odor. It is lighter than water, boils at 320°F . (160°C .) and is a valuable solvent. It is partially oxidized in the air. Many fragrant oils obtained from plants have the same composition and are obtained by distillation. Some of these oils are: lemon, bergamot, coriander, hop, juniper and valerian. They are called essential oils, are mostly lighter than water and freely soluble in alcohol and ether.

Plants furnish us with a number of oxidized turpenes, among which are the camphors and resins.

Common camphor, $C_{10}H_{16}O$, obtained from the camphor laurel, is a white crystalline solid, volatile at ordinary temperatures. It is slightly soluble in water, and freely in alcohol and ether.

Resins include a large group, of which many are true acids and form salts constituting resin soaps.

Common Rosin is the residue from the preparation of oil of turpentine. It is a mixture of two acids.

Lac, Copal and Mastich are familiar members of the group. As a class the resins are easily fusible but not volatile, insoluble in water but soluble in alcohol, which solution constitutes varnish. Gum-resins are simply mixtures of resin and gum; oleo-resins, mixtures of resin and volatile oil; balsams contain benzoic or cinnamic acid.

Caoutchouc and Gutta percha are turpenes found in the juices of plants. They are insoluble in water, but in the plant are usually in suspension, very finely divided so as to make a milky liquid, called an emulsion. Caoutchouc is elastic; gutta percha is not. Both are capable of combining with sulphur to form peculiar and valuable compounds. The process is called vulcanizing.

BENZENES. These begin with C_6H_6 , benzene, benzole or coal-tar naphtha, which is the important member.

Benzole or benzene is obtained by the fractional distillation of coal-tar and by the distillation of benzoic acid with lime. It is a colorless, volatile liquid which solidifies at a little above the temperature of melting ice. Its solvent powers are very high and it is largely used for such purpose. Benzene must not be confounded with the mixture of hydrocarbons obtained by distilling coal oil, and known as benzine.

The other series of hydrocarbons include but few important bodies, of which may be mentioned

Naphthalene, $C_{10}H_8$, occurring in coal-tar. It is a white crystalline solid melting at $176^\circ F.$ ($80^\circ C.$) and slightly soluble in water.

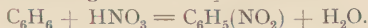
Anthracene, $C_{14}H_{10}$ is obtained from coal-tar. it is especially valuable as a source of alizarine, the coloring matter of madder. When pure it is a colorless solid, subliming at $212^\circ F.$ ($100^\circ C.$), insoluble in water. By direct oxidation anthracene yields anthraquinone.

DERIVATIVES OF THE HYDROCARBONS. With chlorine a large number of substitution compounds may be formed. Those from

marsh gas have been given on page 74. The other paraffins, the olefins and turpenes give similar bodies. Some combine directly with chlorine. Thus olefiant gas gives us $C_2H_4Cl_2$; benzene gives us $C_6H_6Cl_6$. With an oxidizing agent a variety of products is obtained. With nitric acid we have sometimes simple oxidation, sometimes substitution of NO_2 for H. The following is a description of the more important results of these actions.

Chloroform, $CHCl_3$, can be obtained by the action of chlorine on marsh gas, but the commercial article is made by distilling alcohol with bleaching powder. The reaction is not well understood, and the product is at first quite impure. When pure it is a colorless fragrant liquid, very volatile, not easy to burn, insoluble in water and much heavier than that liquid. It boils at $142^\circ F.$ ($61^\circ C.$) It has high solvent powers and is a valuable anæsthetic.

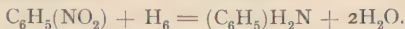
Nitrobenzole, oil of myrbane, $C_6H_5(NO_2)$, is obtained by the action of strong nitric acid upon benzene



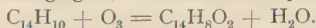
The product is a yellow oily liquid smelling like bitter almonds, insoluble in water and poisonous. It is important because by the action of nascent hydrogen it yields aniline from which many brilliant colors are made.

Nitrobenzole.

Aniline.



Anthraquinone, $C_{14}H_8O_2$, is obtained by the action of a powerful oxidizing agent, as chromic acid, upon anthracene.



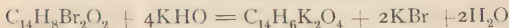
Anthraquinone is used as a source of artificial madder color. The process is, (*a*) the anthraquinone is heated with bromine

Dibromanthraquinone.



(*b*) Dibromanthraquinone heated with potash gives potassium alizarate from which the alizarin may be obtained by sulphuric acid.

Potass. Alizarate



NOMENCLATURE. The names of these derivatives are made by attaching to the original body the name of the added substance, with a prefix indicating quantity.

$C_6H_5(NO_2)$ is nitrobenzole; $C_6H_4(NO_2)_2$ is dinitrobenzole. CH_2Cl_2 is dichloromethane. The different series are indicated by different terminations; starting with C_5H_{12} , quintane we have

by successive removal of H_2 quintene, quintine, etc. Those containing an uneven number of hydrogen atoms terminate in yl. We have methyl, CH_3 , methenyl, CH and so on.

Alcohols and derivatives. The alcohols may be regarded as hydrates in which hydrocarbons are acting the part of the metal. Each series of unsaturated hydrocarbons will give us a series of alcohols, but it be necessary to describe only a few of the more important, such as

Monatomic alcohols, formed from the methyl series:

Diatomic “ “ “ olefine “

Triatomic “ “ “ methenyl “

The members of the first series are the most important.

Methyl Alcohol, Wood Spirit, $(CH_3)HO$, Methyl Hydrate, is usually made by distilling wood. The crude material is difficult to purify. Pure methyl alcohol is colorless, and of pleasant odor. It boils at $152^\circ F.$ ($66.5^\circ C.$) and its effects on the animal system appear to be less severe and more transient than those of common alcohol. The methylated spirit of English chemists is a mixture of 90 parts common alcohol with 10 parts methyl alcohol.

Ethyl Alcohol, Common Alcohol, Spirit of Wine, $(C_2H_5)HO$, Ethyl Hydrate, is produced in the vinous fermentation of sugar; alcohol and carbonic anhydride being chiefly formed; it can also be prepared artificially. On the large scale the sprouted grain called malt is generally used. The general nature of fermentation will be explained in connection with the sugars. The fermented spirit is concentrated by distillation, having a lower boiling point than water, but the strongest spirit thus prepared contains 10 per cent. of water. To withdraw all the water, it is necessary to distill with quicklime, by which *absolute alcohol* is formed. This is very inflammable, greedily absorbs moisture, and mixes with water in all proportions.

Proof-spirit contains 50.8 parts by weight of alcohol to 49.2 of water, and has a specific gravity of 0.920. Commercial alcohol is a colorless volatile liquid of which the properties, effects and uses are well known. The strongest spirit ordinarily furnished is about 95 per cent. and boils at about $180^\circ F.$ ($81^\circ C.$) Alcohol is contained in wine, beer and spirits; certain essential oils, sugar, or extracts, being employed as flavoring agents. Whisky, brandy, and other spirits contain from 40 to 50 per cent. of alcohol; wines, from 17 (port and madeira) to 7 or 8 (hock and light clarets) per cent.; porter and strong ale contain from 6 to 8 per cent.; lager beer about 4 per cent. The effervescence of fermented liquids is due to the carbon dioxide which is produced with the alcohol, thus:

Glucose,

Alcohol

 $C_6H_{12}O_6$ breaks up into $2C_2H_6O + 2CO_2$;

The carbon dioxide is retained by bottling the liquid before the fermentation is over.

Amyl Alcohol, Fusel, Oil, $(C_5H_{11})HO$. Amyl hydrate is a by-product in fermentation, and is found in raw spirits and new liquors. When pure it is a colorless oily liquid, with a peculiar odor, a hot and acid taste, and decidedly poisonous in its action.

The alcohols derived from the higher radicles are mostly wax-like.

Ether, Ethyl Oxide, $(C_2H_5)_2O$, is made by the action of dehydrating agents, especially sulphuric acid, upon alcohol. It appears that acid ethylsulphate is first formed and then decomposed.

Alcohol,

Acid ethylsulphate,



Another molecule of alcohol is then acted upon, thus :

Ether.

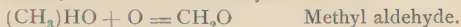


Ether is a colorless, very volatile liquid of distinct odor, boiling at $98^\circ F.$ ($37^\circ C.$) Specific gravity 0.723. Its vapor is inflammable and very heavy. It is a solvent for fats, fixed and volatile oils, resins and many other proximate principles. Methyl and amyl ethers are also known.

Compound Ethers. The alcohol radicles replace the hydrogen of acids and form bodies called compound ethers. It will be sufficient to enumerate a few of the most important. Many of these compound ethers are fragrant and are used as flavoring materials, or as substitutes for fruit essences.

Ethyl Bromide, $(C_2H_5)Br$, is used in medicine. Amyl, acetate, butyrate and valerate are used as flavors. Amyl nitrate is used in medicine. Ethyl nitrate exists in sweet spirit of nitre.

ALDEHYDES AND ACIDS. When alcohols are oxidized by a limited amount of oxygen, two atoms of hydrogen are removed and no oxygen is added. When oxidized in a free supply of oxygen an atom of oxygen takes the place of the removed hydrogen. The bodies produced in the first case are *aldehydes*, in the second, *acids*. In this way we have



Thus each alcohol may be made to yield an aldehyde and an

acid. The acids are very important and will be described under the title fat acids, many of them existing in common fats.

Sugars. The sugars form an important group, the exact relations of which are not well understood, but they are generally regarded as alcohols or aldehydes derived from complex radicles. The group includes sugars proper, also gums, starches and wood fibre. A remarkable similarity of composition, and convertibility into one another by simple means, is to be noted. The most important point in regard to their composition is that they contain oxygen and hydrogen in the proportion to form water. They are divided into three classes.

1. SUCROSES	2. GLUCOSES	3. AMYLOSES (Starch
(Sugars Proper)	(Grape Sugars)	and Woody Fibre)
$C_{12}H_{22}O_{11}$	$C_6H_{12}O_6$	$C_6H_{10}O_5$
Sucrose (Cane	Dextrose (Grape	Starch
Sugar)	Sugar)	Dextrin
Lactose (Milk	Lævulose (Fruit	Inulin
Sugar)	Sugar)	Gum
		Cellulose.

Sucrose or Cane Sugar, $C_{12}H_{22}O_{11}$, exists in certain plants especially sugar-cane and beetroot. It is extracted by pressure; the liquid is then boiled down carefully, the raw product decolorized by animal charcoal and finally crystallized. It is soluble in about twice its weight of cold water. When heated to about 420° F. (216° C.) a **caramel** is formed.

Lactose or Milk Sugar, is only found in milk. It is converted by dilute acids into a peculiar variety of glucose, and in the presence of cheese undergoes lactic fermentation.

Glucose or Grape Sugar, $C_6H_{12}O_6$, is found in manna and honey, and many kinds of fruits. It is a normal constituent of the blood, and is excreted in considerable amount in the disease called diabetes mellitus.

Glucose presents two modifications, dextrose and lævulose, distinguished by their action on light. Dextrose may be obtained by boiling starch with dilute sulphuric acid, adding chalk, and evaporating the liquid. It is soluble in dilute alcohol, but is not nearly so sweet as sucrose. When a solution of cane sugar is boiled with dilute acids, a mixture of dextrose and lævulose is formed, called **inverted sugar**. Lævulose does not crystallize and is sweeter than dextrose.

Fermentation is a peculiar form of decomposition. Organic bodies undergo fermentation under the influence of complicated substances termed ferments, giving products differing according

to the nature of the fermented body and of the ferment. This process appears to depend upon the presence and growth of certain living organisms. Most of these are vegetable growths of a low kind.

The circumstances necessary for the growth of the ferment are : (a) Proper food, especially the ammoniacal salts and alkaline phosphates. These are contained in the albuminous matter generally present in the liquid about to be fermented. (b) A temperature from 60° to 100° F. (20° to 40° C.), since at other temperatures the vitality of the ferment is destroyed.

Spontaneous Fermentation takes place in wine, beer, milk and urine, and is due to the sporules or seeds of living bodies floating in the air. These, dropping into the liquid, propagate themselves, and during the act of growing evolve the products of fermentation. Circumstances destructive or preventive of fermentation are : if the liquid be left in contact only with air which has been passed through a red-hot platinum tube, or if the air be simply filtered by passing it through cotton, and the sporules thus prevented from entering. Thus fermentable liquids may be preserved for any length of time without undergoing change.

The principal forms of fermentation are five in number :

1. The Vinous, producing chiefly Alcohol, C_2H_6O , and Carbonic Anhydride, CO_2 .
2. The Acetous, producing chiefly Acetic Acid, $C_2H_4O_2$.
3. " Lactic " " Lactic Acid, $C_3H_6O_3$.
4. " Butyric " " Butyric Acid, $C_4H_8O_2$.
5. " Mucous " " Gum and Mannite,

Dextrin, $C_6H_{10}O_5$. This substance is also known as British gum, and may be obtained by heating starch to about 320° F. (160° C.). The change is much more speedily effected by the addition of a little hydrochloric or nitric acid. Dextrin, together with dextrose, is formed when malt extract acts upon starch. It is insoluble in alcohol, but very soluble in water; and is used as a mucilage. It is converted into glucose by heating with dilute acids.

Gum Arabic is a natural exudation from many species of *Acacia*. It consists chiefly of arabic acid, $C_{12}H_{20}O_{10}$, united with calcium and potassium.

Starch, $C_6H_{10}O_5$, occurs in many plants. It is a white powder, which is made up of granules of various sizes, having a definite organized structure. These granules are not soluble in cold water, ether or alcohol, but if heated with water to about 160° F. (72° C.), they swell and break up, yielding a thick

mass termed starch paste. Upon boiling this mass with more water, the particles are reduced to such a fine state of division that they will pass through a filter, and when the boiling is continued for sometime the solution becomes clear, and the starch soluble. The test for starch is the formation of a deep-blue color with free iodine.

Starch exists in the seeds of grasses, associated with an albuminous substance, diastase, which has the power to transform the starch into glucose. When the seed germinates this transformation begins, and if the germination be interrupted before the sugar begins to undergo further change we have malt, which is simply sprouted grain, especially barley. When malt is steeped in water and yeast added, the fermentation of the sugar begins. Dilute sulphuric acid acts like diastase.

Cellulose, $C_6H_{10}O_5$, is the colorless material of woody fibre. It is obtained nearly pure by boiling cotton with alkali. Cellulose is a white substance which dissolves in an ammoniacal solution of cupric oxide, but is insoluble in water, ether, or alcohol. Strong sulphuric acid converts it either into a soluble substance like dextrin, or into an insoluble substance, giving a blue color with iodine.

By dipping sheets of paper into strong sulphuric acid, *parchment paper* is obtained.

Gun-cotton. When cotton is put into a mixture of equal volumes of strong nitric and sulphuric acids, no apparent change occurs; but after drying it is found to be exceedingly inflammable. A substitution product is here formed, termed trinitro-cellulose, in which NO_2 replaces hydrogen, $C_6H_7(NO_2)_3O_5$.

Collodion is formed by dissolving certain kinds of gun-cotton in a mixture of ether and alcohol. It is much employed in photography and surgery.

Mannite, $C_6H_8(OH)_6$, or $C_6H_{14}O_6$, is a sugar-like substance contained in manna, an exudation from many species of ash. When acted upon by nitric acid it forms a compound termed nitro-mannite, or mannite in which six atoms of hydrogen are replaced by NO_2 .

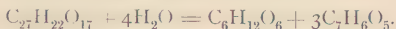
Glucosides. These occur in many plants, and on decomposition give rise to a glucose, together with other substances. Only a few need be mentioned.

Amygdalin, $C_{20}H_{27}NO_{11} + 3H_2O$, is found in bitter almonds. It forms white soluble crystals. By the action of water (H_2O) upon the amygdalin and the albuminous substance termed Synaptase or Emulsin, contained in the bruised bitter almond, a species of fermentation is set up in which hydride of benzoyl (C_7H_6O), hydrocyanic acid (HCN), and glucose ($C_6H_{12}O_6$) are formed. Thus—



Salicin, $\text{C}_{13}\text{H}_{18}\text{O}_7$, is found in the pith of the poplar and willow. It crystallizes in white brilliant needles. It is soluble in water and alcohol, the solution having a strong bitter taste. It yields saligenin and glucose.

Tannin, or Tannic Acid, $\text{C}_{27}\text{H}_{22}\text{O}_{17}$, is widely diffused in certain parts of plants. It forms an insoluble substance with gelatin, and yields a bluish-black color (ink) with ferric salts. The largest quantities of tannic acid are found in the gall-nuts, excrescences formed by the punctures of an insect. Tannin is uncrystallizable, insoluble in pure ether, but dissolved by water and alcohol. When exposed to the air or treated with dilute acid, glucose and gallic acid are formed.



The sugars and starches are reducing agents, but not very energetic. The action is generally increased by the presence of strong alkalies. The tests for them are mostly dependent upon their reducing action upon the salts of copper, silver, bismuth and mercury. Glucose is especially active.

Moore's test depends on the fact that a solution of sugar becomes darker on being boiled with caustic alkali.

Froehde's test is performed by adding to the suspected solution a few drops of copper sulphate, and a considerable amount of caustic soda or potassa. If sugar be present the application of heat will cause the precipitation of orange-red cuprous oxide, Cu_2O .

Boettger's test is performed in the same way, substituting bi-muth oxynitrate for copper sulphate. A black precipitate of metallic bismuth is formed. The most accurate test is by fermentation, and recognition of the alcohol.

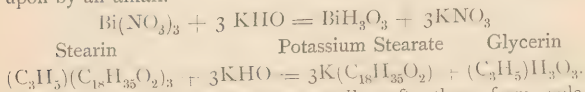
DIATOMIC ALCOHOLS, GLYCOLS. These are not important. Ethene glycol, $(\text{C}_2\text{H}_4)\text{H}_2\text{O}_2$, is the best known. It is a thick liquid, sweet and colorless.

TRIATOMIC ALCOHOLS, or GLYCERINS. Of these the best known is glycerin (C_3H_5) H_3O_3 . It is obtained by the decomposition of fats. It does not oxidize nor evaporate in the air, and has a solvent action next to that of water. Heated with strong nitric acid it forms nitro-glycerin, $\text{C}_3\text{H}_5(\text{NO}_2)_3\text{O}_3$, which is a violent explosive, and when mixed with sand or other material constitutes **dynamite**.

Fats and Fixed Oils. These are compound ethers, mostly derived from glyceryl, (C_3H_5) , a triad, and therefore, like bismuth, requiring three molecules of a monobasic acid

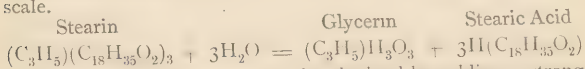
to form a normal salt. As found in the tissues of plants and animals they are mixtures of several distinct bodies which can be separated by difference of melting point, solubility, etc. The oils are merely fats with a low melting point. They are usually divided into two classes: drying oils, which absorb oxygen from the air, and become hard and resinous, such as linseed and poppy oil; non-drying oils, which remain fluid, as castor and sperm oil. Many fats and oils undergo partial decomposition in the air, producing a free acid; this is called rancidity.

Each distinct fat is usually called by a name derived from the acid that forms it. Glyceryl oleate is called *olein*; glyceryl stearate is called *stearin*. When caustic alkali is added to a fat, decomposition takes place, a salt of the alkali-metal is formed constituting a soap, and glycerin is produced. The reaction is analogous to that which occurs when bismuth nitrate is acted upon by an alkali.



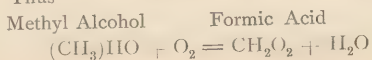
Soaps produced by potassa are usually soft; those from soda hard; those made from other oxides are mostly insoluble in water. This latter fact explains the curdling action of hard water. The calcium and magnesium salts produce insoluble soaps. When soaps of the alkali-metals are treated with cold water they decompose into acid salt, which precipitates and makes the soap-suds, and an oxysalt, which dissolves and gives the cleansing action.

A decomposition of the fats may be produced by the action of super-heated steam. In this case glycerin and free acids are formed and can be distilled off. The process is used on the large scale.



The acids of the fats may also be obtained by adding a strong acid to ordinary soaps.

Fat Acids. These form an extensive and important set of compounds. Most of them bear a direct relation to the alcohols, from which they may be produced by the substitution of one atom of oxygen for two atoms of hydrogen in the alcohol. The removed hydrogen is converted into water by another atom of oxygen. Thus



The acids so formed from the monatomic alcohols are monobasic, that is, only one atom of hydrogen is capable of being replaced by a metal. They are therefore usually written with a single atom of hydrogen separate from the rest; formic acid, for instance, is written HCHO_2 . The following are some of these acids and their corresponding alcohols:

Alcohol.	Acid.
Methyl Alcohol, $(\text{CH}_3)\text{HO}$	yields HCHO_2 , Formic Acid.
Ethyl " $(\text{C}_2\text{H}_5)\text{HO}$	" $\text{HC}_2\text{H}_3\text{O}_2$, Acetic Acid.
Propyl " $(\text{C}_3\text{H}_7)\text{HO}$	" $\text{HC}_3\text{H}_5\text{O}_2$, Propionic Acid.
Butyl " $(\text{C}_4\text{H}_9)\text{HO}$	" $\text{HC}_4\text{H}_7\text{O}_2$, Butyric Acid.
Amyl " $(\text{C}_5\text{H}_{11})\text{HO}$	" $\text{HC}_5\text{H}_9\text{O}_2$, Valeric Acid.

Only a few of this series of acids will need description.

FORMIC ACID, HCHO_2 , exists as a secretion in some stinging animals and plants. It is a corrosive volatile liquid and acts as a reducing agent.

ACETIC ACID, $\text{HC}_2\text{H}_3\text{O}_2$. This, in the dilute state, constitutes vinegar which contains about 5 per cent. of the acid, and is usually made by oxidizing dilute alcohol in the presence of a ferment. Acetic acid is also produced in the distillation of wood. When pure, it is a colorless corrosive liquid, solidifying at 62.6°F . (17°C), and boiling at 246°F . (119°C .) This is glacial acetic acid. The more dilute forms are less active, and in vinegar, its effects are quite mild.

Acetates. Many of these are important. **Potassium and sodium acetates** are very deliquescent.

Lead acetate, $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$, sugar of lead, made by dissolving lead oxide in acetic acid, forms white crystals, soluble in water. The solution is capable of taking up more lead oxide, and forming an oxyacetate, known as Goulard's extract.

Cupric oxyacetate is a compound of copper acetate with copper hydrate, and is known as *verdigris*.

VALERIC ACID. This is obtained by the action of oxidizing agents on amyl alcohol. It also exists in some plants. It is an oily liquid of disagreeable odor. Several of its salts are in extensive use in medicine.

The higher members of this monobasic series are generally insoluble in water, and oily or fatty in appearance.

Acids from Glycols. Two series of these exist—one derived by the replacement of two atoms of hydrogen by one atom of oxygen, and the other by the replacement of four atoms of hydrogen by two atoms of oxygen. The first is the lactic acid series; the second, the oxalic acid series.

The following will show the relation between the two groups :

Glycol.	1st oxidation.	2nd oxidation.
$C_2H_6O_2$	$C_2H_4O_3$, Glycollic	$C_2H_2O_4$, Oxalic.
$C_3H_8O_2$	$C_3H_6O_3$, Lactic	$C_3H_4O_4$, Malonic.
$C_4H_{10}O_2$	$C_4H_8O_3$, Oxybutyric	$C_4H_6O_4$, Succinic.

Each series is homologous, and carbonic acid CH_2O_3 may be regarded as the first member of the first series.

Only a few of these acids need description.

Lactic acid, $HC_3H_5O_3$, is contained in sour milk, and is formed from sugar during the so-called lactic fermentation. It may be prepared artificially. It is a syrupy liquid which forms soluble salts.

Oxalic acid, $H_2C_2O_4$, exists ready formed in the juices of some plants, as sour sorrel and garden rhubarb. It is made artificially by the action of nitric acid on sugar, or of caustic alkali on sawdust. It is sometimes a result of disease in the animal system. It occurs in colorless crystals which contain $2H_2O$. It is soluble in water, very sour and highly poisonous. Many of its salts are insoluble in water, especially calcium oxalate. Its decomposition with sulphuric acid is a convenient source of carbon monoxide (page 52.) With the monad metals two oxalates are formed. Acid potassium oxalate, KHC_2O_4 , occurs in the leaves of plants, and is called salt of sorrel. Calcium oxalate, CaC_2O_4 , is occasionally formed in the kidney. Ammonium oxalate, $Am_2C_2O_4$, is used as a test.

Succinic acid is produced from amber, a fossil resin, and exists in a few plants. It can also be made by the action of nitric acid upon most fats.

Other Organic Acids. A great many organic acids are known, which belong to different incomplete series which it will be unnecessary to enumerate. It will be sufficient to mention some important acids, and indicate their sources and properties.

Malic acid, $H_2C_4H_4O_5$, occurs in the juices of many plants, as apples, pears, mountain ash berries and tobacco leaves. It may be made artificially. It is crystalline, sour, soluble in water and alcohol. Its salts have very little importance.

Tartaric acid, $H_2C_4H_4O_6$. This is found in many plants but especially in grapes, where it exists as acid potassium tartrate, $KHC_4H_4O_6$. This salt is somewhat soluble in water, but scarcely soluble in dilute alcohol, and hence in the manufacture of wine, as the fermentation advances, the quantity of alcohol increases, and the acid potassium tartrate deposits as a red mass called argols. This being dissolved in hot water and crystallized, gives cream of tartar, (called very wrongly by some writers, *cremor tartar*).

Tartaric acid is a crystalline body soluble in water and very sour. Its solution in water develops a fungous growth and decomposes. Several varieties of the acid are known.

Acid potassium tartrate is a white crystalline body very sour and not very soluble in cold water. It is used in effervescing powders.

Potassium tartrate, $K_2C_4H_4O_6$, is called soluble tartar.

Sodio-potassium tartrate, $NaKC_4H_4O_6$, is known as Rochelle salt.

Tartar emetic is described on page 87.

Citric acid, $H_3C_6H_5O_7$, is the acid of lemons and oranges, and is also found in some other fruits. It is a crystalline body, very sour, and easily soluble in water.

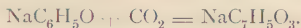
Oleic acid, $HC_{18}H_{33}O_2$, exists in most natural fats and non-drying oils. It is solid at 57° F. (14° C.) Above this temperature it is a clear liquid, lighter than water and insoluble in it, but soluble in alcohol and ether. Crude oleic acid, made by the decomposition of fats by steam, as mentioned on page 148, is used in soap-making under the name of red-oil.

Benzoic acid, $HC_7H_5O_2$, is found in many resins, but chiefly in benzoin, by heating which, the acid sublimes in white, pearly plates. Most of the salts formed by benzoic acid are soluble.

Salicylic acid, $HC_7H_5O_3$, is formed by the oxidation of salicin, the bitter principle of the willow. It may be obtained by acting upon phenol (C_6H_5HO) with carbon dioxide (CO_2) sodium (Na). Thus

Sodium Phenylate

Sodium Salicylate



The acid also occurs in the oil of winter green, methyl salicylate, $(CH_3)C_7H_5O_3$.

Carbolic acid, HC_6H_5O , is hardly a true acid, but is so called from a power of taking up bases and forming bodies like salts. It is preferably called phenol or phenylic alcohol. It exists in coal-tar, and can also be made by several processes. It forms colorless crystals, is very deliquescent and soluble in water. It melts at 93° F. (34° C.), and boils at 370° F. (187° C.) It has a peculiar odor, much like that of kreasote, which is a somewhat similar body obtained from wood-tar. Phenol, salicylic acid and benzoic acid are used as antiseptics, as they possess the power to retard the development of many forms of microscopic life.

Picric acid, or **tri-nitro phenol**, $C_6H_3(NO_2)_3O$, is a yellow crystalline substance, very soluble in water, and is formed when nitric acid acts upon phenol. It may also be obtained when nitric acid acts upon many other substances, as the skin, for instance. In the arts it is employed as a dye for silk and wool, It is monobasic, and some of its salts are explosive.

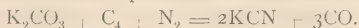
Gallic acid, $HC_7H_5O_5$, is obtained from tannin. It is used as an astringent. By heat it yields **pyrogallin**, $C_6H_6O_3$, sometimes called **pyrogallic acid**, which has the properties of absorbing oxygen rapidly when mixed with a base, and is used as a test for oxygen, and as a reducing agent for photography.

Compounds containing Nitrogen. These include the most complex organic bodies. Carbon and nitrogen form the cyanogen series. Carbon, hydrogen and nitrogen form a group including many purely artificial products having alkaline characters. Carbon, hydrogen, nitrogen and oxygen form a group including most of the principles upon which the medicinal and poisonous virtues of plants depend; this group includes also many of the materials forming the tissues of animals and plants.

Cyanogen and derivatives. Cyanogen, CN , is produced when organic matter containing nitrogen, such as leather scraps, is heated with an alkali, especially in the presence of iron. Cyanogen is a gas, but is of no importance in the free state. It is often represented by the symbol Cy (see page 57).

Hydrogen cyanide HCN . **Hydrocyanic** or **prussic acid**. This body is not an acid, although often called so. It is produced by the action of acids upon cyanides. It is an easily decomposed liquid of a rather pleasant odor, and intensely poisonous, a drop or two causing death almost instantaneously. It is used in medicine in a much diluted form. Hydrogen cyanide is produced in the fermentation of some glucosides.

Potassium cyanide, KCN , can be made by passing nitrogen over a mixture of potassium carbonate and charcoal :



It is usually made from **potassium ferrocyanide** (q. v.) It is in white fusible deliquescent masses, and is a violent poison probably because it easily furnishes HCN by the action of even feeble acids. It is extensively used in photography in order to dissolve the unaltered salts of silver, a double salt of potassium and silver resulting ($KCN + AgCN$). It is also employed in silver and gold plating.

The cyanides have a strong tendency to form double salts, and several of these are of much importance.

Potassium ferrous cyanide, $\text{FeCy}_2 + 4\text{KCy}$, usually called potassium ferrocyanide and written $\text{K}_4\text{Cy}_6\text{Fe}$, is formed by heating a mixture of nitrogenous organic matter, iron scraps and potassium carbonate, and treating the mass with water. The salt forms large lemon yellow crystals which are not poisonous. It is much used in dyeing, under the name of yellow prussiate of potash. Oxidizing agents convert the ferrous cyanide into ferric, and produce a body called

Potassium ferric cyanide, $\text{Fe}_2\text{Cy}_6 + 6\text{KCy}$, commonly called potassium ferricyanide or red prussiate of potash. It forms large ruby red crystals, soluble in water and used in dyeing. From these double cyanides others may be obtained, especially with iron. When ferrocyanide is added to a ferric salt, a blue precipitate of Prussian blue is formed; when ferricyanide is added to a ferrous salt, Turnbull's blue is formed.

Cyanogen forms several acids which have the same percentage composition, but different formulæ:

HCNO	cyanic.
$\text{H}_2\text{C}_2\text{N}_2\text{O}_2$	fulminic.
$\text{H}_3\text{C}_3\text{N}_3\text{O}_3$	cyanuric

Such bodies form a polymeric series. We have also a sulphur acid, HCNS , which forms sulphocyanates.

Artificial organic bases. The compounds containing carbon hydrogen and nitrogen are mostly artificial products. They are obtained by the substitution of organic radicles for the hydrogen of amine NH_3 , or of ammonium salts. The chemistry of these bodies is so extensive that in this work little more can be done than to indicate the nature of the them. One or more atoms of hydrogen in NH_3 or NH_4 may be replaced, and different radicals may be introduced at the same time. The nitrogen may be replaced by the other members of its group, and thus the changes are almost endless. Bodies so obtained are called substitution ammonias. They are more or less alkaline, and generally form salts with acids. They name according to the character and amount of the substituting body; thus:

Amine.	Phosphine	Ammonium iodide.
H_3N	H_3P	NH_4I
$(\text{CH}_3)\text{H}_2\text{N}$, methylamine.		
$(\text{CH}_3)(\text{C}_2\text{H}_5)_3\text{I}$, methyl-triethyl-ammonium iodide.		
$(\text{C}_2\text{H}_5)_3\text{N}$, triethylamine.		
$(\text{CH}_3)_4\text{NHO}$, tetra-methyl-ammonium hydrate.		
$(\text{CH}_3)(\text{C}_2\text{H}_5)(\text{C}_5\text{H}_{11})(\text{C}_6\text{H}_5)\text{NHO}$, methyl-ethyl-amyl-phenyl-ammonium hydrate.		

Phenylamine, $(C_6H_5)H_2N$ is commonly known as aniline, and is a source of a series of very valuable colors.

Conia, $C_8H_{15}N$, the active principle of water-hemlock; and nicotia, $C_{10}H_{14}N_2$, the active principle of tobacco, are members of this group.

Natural Alkaloids. The medicinal virtues of plants are sometimes due to resins or volatile oils, but generally to nitrogenous bodies which are strongly basic in character, and evidently related to ammonia, but their exact nature is not made out. They are mostly insoluble in water, but soluble in alcohol. With acids they generally form crystalline salts, and are therefore called alkaloids. It will be sufficient to give a table showing the composition, sources, and important properties of the more important.

Alkaloids.	Sources.	Properties.	Formulae.
Conia	Water Hemlock	Colorless liquid	$C_8H_{15}N$
Nicotia	Tobacco	Oily liquid	$C_{10}H_{14}N_2$
Quinia	Peruvian Bark	White crystals	$C_{20}H_{24}N_2O_2$
Cinchonia	"	" "	$C_{20}H_{24}N_2O$
Morphia	Opium	" "	$C_{17}H_{19}NO_3$
Atropia	Belladonna	" "	$C_{17}H_{23}NO_3$
Veratria	Hellebore	White powder	$C_{32}H_{52}N_2O_8$
Strychnia	Nux Vomica	White crystals	$C_{21}H_{22}N_2O_2$
Caffeina	Coffee and tea	" "	$C_8H_{10}N_4O_2$

Urea, CH_4N_2O , is of animal origin.

The fully developed tissues and organs of animals are of such complicated composition that little information has been obtained in regard to them. They contain other elements besides those mentioned in the preceding compounds; sulphur, phosphorus and iron being the most important.

ATOMIC WEIGHTS & SYMBOLS.

Aluminium	Al	27.5	Mercury	Hg	200
Antimony	Sb	122	Molybdenum	Mo	96
Arsenic	As	75	Nickel	Ni	59
Barium	Ba	137	Niobium	Nb	94
Beryllium	Be	9.3	Nitrogen	N	14
Bismuth	Bi	210	Osmium	Os	199
Boron	B	11	Oxygen	O	16
Bromine	Br	80	Palladium	Pd	106.5
Cadmium	Cd	112	Phosphorus	P	31
Cæsium	Cs	133	Platinum	Pt	196.7
Calcium	Ca	40	Potassium	K	39
Carbon	C	12	Rhodium	Rh	104.3
Cerium	Ce	141	Rubidium	Rb	85.4
Chlorine	Cl	35.4	Ruthenium	Ru	104.4
Chromium	Cr	52.5	Selenium	Se	79.5
Cobalt	Co	59	Silicon	Si	28.5
Copper	Cu	63	Silver	Ag	108
Didymium	Di	147	Sodium	Na	23
Erbium	Er	169	Strontium	Sr	87.5
Fluorine	F	19	Sulphur	S	32
Gallium	Ga	68.9	Tantalum	Ta	182
Gold	Au	196.7	Tellurium	Te	128
Hydrogen	H	1	Thallium	Tl	204
Indium	In	113	Thorium	Th	231.5
Iodine	I	127	Tin	Sn	118
Iridium	Ir	198	Titanium	Ti	48
Iron	Fe	56	Tungsten	W	184
Lanthanum	La	139	Uranium	U	240
Lead	Pb	207	Vanadium	V	51.2
Lithium	Li	7	Yttrium	Y	93
Magnesium	Mg	24	Zinc	Zn	65
Manganese	Mn	55	Zirconium	Zr	90

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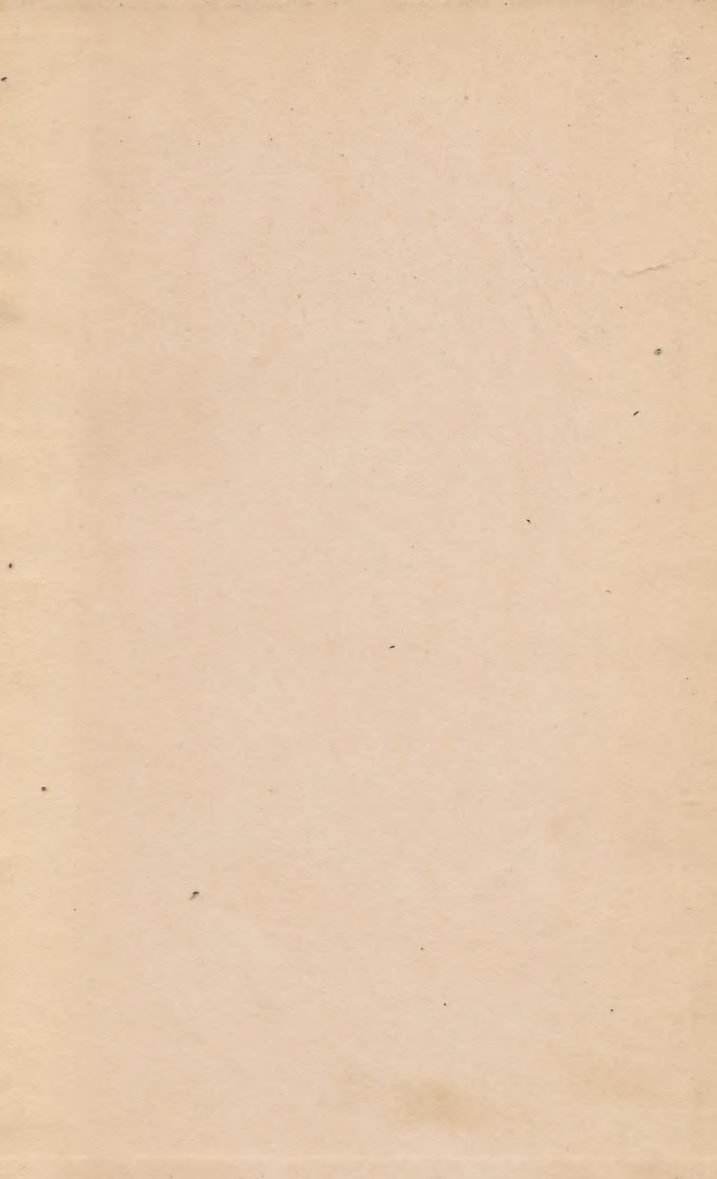
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